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ABSTRACT

This handbook is based on a list of essential topics that should be mastered by the student who subsequently plans to pursue college chemistry. Chapters include: (1) "Introduction" (describing a position paper and the background of the handbook); (2) "Essential General Topics and Objectives"; (3) "Testing Students" (providing sample tests and answers); (4) "The Laboratory" (including a list of laboratory skills, sample tests on laboratory learning, procedures for preparing reagent solutions, and price lists of chemicals and supplies); (5) "Chemical Safety" (describing safety guidelines and checklists with resource materials); (6) "Chemistry Teachers and the Law" (relating accidents in the laboratory); (7) "Demonstrations" (describing the materials, procedures, waste disposal, and questions for 20 demonstrations); (8) "Experiments" (describing the objectives, preparations, sample data, waste disposal and questions for 17 experiments); and (9) "Reference Materials." (YP)

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# STATE OF LOUISIANA
# DEPARTMENT OF EDUCATION

# HANDBOOK FOR HIGH SCHOOL CHEMISTRY TEACHERS

**Bulletin 1839**
**Revised 1988**

**Developed by the Education Committee**
**Mississippi Section**
**American Chemical Society**
**Funded by the National Science Foundation**

SE 051 042

**Wilmer S. Cody**
**Superintendent of Education**

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# HANDBOOK FOR HIGH SCHOOL CHEMISTRY TEACHERS

*Produced by the Education Committee*
*Mississippi Section*
*American Chemical Society*

Editors/Contributing Authors:

John H. Bedenbaugh
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September 1985

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## DISTRIBUTION OF THIS BOOK

It is intended that every person in Mississippi who teaches a high school chemistry course receive a copy of this handbook. If you know of someone teaching even one section of a chemistry course who did not get a copy of this book, please give him/her one of the order forms. The book will be mailed promptly on our receipt of the order form.

We anticipate a major problem in getting copies to each teacher of chemistry in large school systems. Please help us by making copies of the order form and distributing them to those who were not included in our original mailing.

On the other hand, if you receive a copy of this book and are no longer involved in teaching chemistry at your school, please give the book to the current chemistry teacher.

# ACKNOWLEDGEMENTS

The Education Committee of the Mississippi Section of the American Chemical Society wishes to gratefully acknowledge the financial support of the National Science Foundation which made possible the development and production of this book.

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## DISCLAIMER

An intensive effort has been made to minimize any possible hazards associated with the demonstrations and experiments described in this book. A lengthy chapter on laboratory safety is evidence of our concern for the safety of teachers and students who use material from this book. However, lists of hazardous chemicals found in that chapter are not all-inclusive. Furthermore, although warnings of the hazard of using a given chemical appear throughout the book, typically this information does not include all possible hazards associated with handling that chemical.

Obviously it is impossible to foresee each and every problem that could conceivably arise from the use of this book. No laboratory activity should be performed by a teacher or student unless he/she has a full understanding of the chemical or physical phenomena involved, adequate protective equipment, and a "feeling" of confidence and personal security in doing the activity. If there appears to be ambiguity, lack of needed information or lack of clarity in the instructions, DON'T DO THE ACTIVITY. If there seems to be any element of danger or risk associated with doing a demonstration or experiment, DON'T DO THE ACTIVITY. No laboratory activity is important enough to warrant risking the safety of the teacher or of the students.

The authors, editors, and sponsoring agencies accept no moral or legal responsibility for injuries that may result from demonstrations or laboratory activities based on information in this book.

## DUPLICATION PRIVILEGES

Teachers are encouraged to make copies of any part of this publication.

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# CHAPTER 1

# INTRODUCTION

# CHAPTER 1 -- INTRODUCTION

## A POSITION PAPER ON THE TEACHING OF HIGH SCHOOL CHEMISTRY

by The Education Committee of the
Mississippi Section of the American Chemical Society

(This paper represents a consensus opinion of those faculty members primarily responsible for freshman chemistry instruction in the colleges and universities of Mississippi.)

1. We believe that students entering a freshman chemistry course in college should have had at least one year of high school chemistry with a strong laboratory orientation. Our committee has prepared a list of topics that we think should be included in a first-year high school chemistry course. This list matches closely other lists of topics distributed nationally.

2. First-year algebra should be a prerequisite for first-year chemistry in high schools because learning chemistry involves the use of algebraic expressions. Moreover, it is highly desirable to correlate algebra with problems in chemistry and physics to give students a grasp of the utility of algebra.

3. We strongly recommend that high school students who have taken first-year chemistry elect first-year physics rather than second-year chemistry if only one of the two can be taken.

4. It is recommended that a second-year high school chemistry course consist of a study of special topics. Our committee has prepared a suggested list of topics for the second-year course. Neither a year of organic chemistry nor a year of college-level general chemistry is advised for the second high school chemistry course.

5. We discourage attempts to teach advanced placement chemistry in Mississippi high schools at this time. Most of these courses involve only very little laboratory work which is so essential to an understanding of chemistry. Moreover, we are of the opinion that relatively few persons now teaching high school chemistry in Mississippi have the required training to teach a good advanced placement course in chemistry. High school advisors might instead encourage good students to enroll in advanced placement courses in such subjects as English, mathematics, etc. which don't require laboratory work, but only if teachers capable of teaching college-level courses are on the staff.

6. We endorse the "Guidelines and Recommendations for the Preparation and Continuing Education of Secondary School Teachers of Chemistry" prepared and published by the American Chemical Society.

7. We urge administrators and teachers to cooperate in rigidly adhering to a schedule that gives the correct allotted time for academic subjects. Students should not be allowed to miss chemistry classes to participate in any extracurricular activities.

8. We urge teachers who plan to take college-level courses for recertification to enroll in content-oriented rather than methods-oriented courses. We believe that our high school teachers already know how to teach but, like all scientists, must constantly work to stay current in the subject matter.

9. We cannot stress too much the importance of the laboratory experience to a good course in high school chemistry. Not only does it stimulate the student to think and enhance his interest in the subject; it is absolutely fundamental to his basic understanding of chemistry. It is our opinion that, with a few exceptions, meaningful chemistry laboratory study is sadly lacking in Mississippi high schools. We therefore urge

   Teachers: to give more emphasis to laboratory instruction and to prepare themselves through in-service training to better instruct in a laboratory setting.

   Administrators:

   (a) to reduce the number of courses taught by chemistry teachers in order to allow adequate preparation and clean-up time for laboratory instruction.

   (b) to develop flexible scheduling to accommodate laboratory instruction. Blocks of time longer than the usual 50-minute periods are necessary for good laboratory instruction.

   (c) to budget considerably more money to support good-quality laboratory instruction. Most high schools in Mississippi have a minuscule budget for laboratory equipment and supplies. There must be a large increase in most budgets if quality laboratory programs are to be developed.

   (d) to make the necessary arrangements to enable teachers to participate in in-service training in laboratory instruction with minimum personal inconvenience. Teachers should be encouraged to participate by use of appropriate incentives.

# WHY THIS BOOK WAS WRITTEN

Colleges and universities today expect students to have studied chemistry in high school. If these students are to do well in college chemistry, they should have earned good grades in a quality chemistry course in high school.

It became obvious in the early 1970s that freshmen by the hundreds were appearing on Mississippi college campuses having no chemistry in their high school backgrounds or having taken courses so poor that they could not pass simple placement examinations. Three choices were open to the colleges and universities: (1) they could put these students into the typically rigorous freshman courses in general chemistry on a "sink or swim" basis, (2) the rigor and quality of the freshman courses could be lowered to accommodate these underprepared students, or (3) remedial sub-freshman courses in chemistry could be developed and offered on college campuses as prerequisite courses for the underprepared student wanting to enter the general chemistry sequence. Option 1 seems unfair to the good student with a poor background; option 2 was never even considered. Therefore, there began a proliferation of remedial chemistry courses at colleges and universities across the state. Today eight post-secondary institutions (including our largest universities as well as our junior colleges) offer such courses. Hundreds of students are enrolled yearly in these courses! Parents who pay taxes for a good high school education for their children must now also pay college tuition for their children to take high school chemistry on a college campus because these remedial courses must be successfully passed by the underprepared freshman before the standard college-level courses can be taken with any expectation of success.

Those college and university teachers most closely associated with the general chemistry programs at their institutions, seeing no improvement in the quality of high school chemistry instruction over twelve years, felt compelled to try to help in some way. Uniting under the banner of the Education Committee of the Mississippi Section of the American Chemical Society, they have contacted scores of high school chemistry teachers in Mississippi to try to ascertain the problem and decide how the Education Committee can help to improve the quality of high school chemistry. It soon became obvious that the shortage of high school chemistry teachers has forced numerous school systems to employ as chemistry teachers many persons whose academic backgrounds are in areas other than chemistry. While these persons may have excellent teaching skills, the fact is that they are deficient in a knowledge of chemistry. Frequently this results in classes in which the main activity is merely reading the text. Some teachers voiced the fact that they are literally afraid to perform demonstrations and conduct laboratories because of a lack of knowledge of the chemistry involved and, therefore, a lack of confidence in their ability to use successfully these effective methods of teaching.

The Education Committee concluded that massive state-wide retraining of chemistry teachers in the subject matter is urgently needed. Since such a project will take years, it appeared that the best short-term help might be a handbook for high school chemistry teachers to assist those with a limited background to learn more about the subject and gain the confidence needed to present a better course in chemistry. Subject matter was carefully selected to make the handbook also useful to well-prepared, experienced teachers of high school chemistry. We see the book as a source of many types of important information that is not found anywhere else in one volume.

## HOW THIS BOOK WAS WRITTEN

Beginning in 1983 twelve members of the Education Committee of the Mississippi Section of the American Chemical Society, representing all of the state universities and the larger private colleges in Mississippi, held meetings as time and finances permitted to develop materials to assist high school chemistry teachers.

These materials were so well received by high school chemistry teachers that the authors envisioned an accumulation of materials in the form of a handbook for high school chemistry teachers. A Materials Development Grant from the National Science Foundation in 1984 gave substance to the dreams of the group. Three of the most outstanding high school chemistry teachers in Mississippi were selected for inclusion in the materials-writing group which became known as the Action Group of the Education Committee of the Mississippi Section of the American Chemical Society.

The Action Group, worked independently much of the time, but also met in the spring of 1985 in two intensive one-week workshops to write materials. A number of consultants made brief visits to these workshops to give advice and make recommendations. Members of the Action Group spent hundreds of hours selecting and personally developing demonstrations and experiments which will allow a high school chemistry teacher with a minimum background in chemistry to begin a modest laboratory program with confidence.

Most of the other features of this book are derived from the collective wisdom and teaching experience of the Action Group. An exception is the chapter on safety which was largely written by the high school and junior college teachers who were enrolled in a course, Chemical Safety in the Academic Laboratory, taught by one of the editors (JHB) during the summer of 1985.

In the summer of 1985 the two editor-authors completed the writing, organizing, and editing of the contents of this handbook.

# THE DESIGN AND SUGGESTED USE OF THIS BOOK

This handbook is organized around a minimum list of essential topics that should be mastered by the student who subsequently plans to pursue college chemistry. These topics are keyed to pages in several of the more popular current high school chemistry textbooks. A suggested allocation of class time to each topic is given.

Acknowledging the trend to "teach by objectives", a minimum list of learning objectives has been carefully developed and keyed to the essential topics.

Considerable attention is given to the subject of testing the students. Numerous sample tests are included along with suggestions for taking some of the drudgery out of the testing process.

Because the Education Committee is convinced that laboratory experiences are critically important to a good high school chemistry course, much of this handbook has been devoted to laboratory instruction. There are approximately 135 pages of detailed procedures and guide sheets for demonstrations and experiments that are all keyed to the essential topics. It is hoped that this material will help the teacher to involve students in seeing and doing chemistry rather than merely reading about it. All experiments have been designed to: (1) be done in 50 minutes, (2) use low cost materials, (3) be safe, and (4) make a worthwhile scientific point. Action Group members have personally conducted the experiments and, in many cases, have supervised use of them with classes of high school students.

The chapter on safety should be read before attempting to do any of the demonstrations and experiments described in this book. One of the concerns of the authors of the safety chapter is the personal liability assumed by every chemistry teacher who conducts demonstrations and experiments. Reflecting this concern is a brief section of the book devoted to suggestions for minimizing personal liability.

We urge the reader to review carefully the Table of Contents to learn of other useful features to be found in this book.

We want this book be used by teachers. To that end, it is distributed in a 3-hole punched, non-bound format for use in a 3-hole notebook binder. Thus pages of the book may be removed and copied to obtain multiple copies for any instructional purpose conceived by the teacher. Notes made by the teacher can be integrated easily into the handbook to make it a more useful personalized reference.

# A WORD TO THE TEACHER

This book does not represent an attempt on the part of the Education Committee of the Mississippi Section of the American Chemical Society to dictate to a high school chemistry teacher what to teach or how to teach. We present in a spirit of helpfulness a minimum number of topics and learning objectives which a college-bound student must master in high school chemistry if he/she is to be well prepared for a general chemistry course in college.

Expansion of these minimum topics, inclusion of other topics, and the "personalizing" of a chemistry course to reflect the personality and innovativeness of the individual teacher are all strongly encouraged.

Many of the demonstrations and experiments described in this book are by design quite elementary in concept and simple in execution. Chemistry teachers who already have an active and successful program of demonstrations and experiments may choose to use only a few, or perhaps none, of those presented in this book. On the other hand, chemistry teachers who, for various reasons, now have little or no laboratory activity in their courses should find this book very helpful in beginning a laboratory program that can be expanded with time.

# CHAPTER 2

# ESSENTIAL GENERAL TOPICS AND OBJECTIVES

# CHAPTER 2 -- ESSENTIAL GENERAL TOPICS AND OBJECTIVES

## DEVELOPMENT OF THE ESSENTIAL GENERAL TOPICS LIST

The information most wanted by high school teachers from the Education Committee has been a list of high school chemistry topics the mastery of which is considered essential to success in college chemistry. Some teachers have been teaching esoteric topics in organic chemistry or personally favorite topics while giving short shrift to basic topics that are absolutely fundamental to the understanding of topics generally covered in college chemistry.

After considerable discussion and deliberation, the Education Committee developed a list of essential topics that should be taught well in high school chemistry courses if a student is to succeed subsequently in college chemistry. Research shows that this list is very similar to essential topics lists developed independently by other groups in other areas of the United States.

This Essential General Topics List is the focal point of this handbook and the cornerstone around which the rest of the book is constructed.

## ESSENTIAL GENERAL TOPICS FOR HIGH SCHOOL CHEMISTRY

(With a Proposed Sequence of Topics and Allocation of Time to Each Topic)

| | Topic | Class Periods |
|---|---|---|
| I. | Introduction (Nature, Science, Chemistry) | 1 |
| II. | Metric System and Density | 6 |
| III. | Scientific Notation and Significant Figures | 4 |
| IV. | Naming, Symbols, and Formulas | 7 |
| V. | Equations | 6 |
| VI. | Stoichiometry and the Mole Concept | 10 |
| VII. | Descriptive Chemistry | 10 |
| VIII. | Energy and Chemical Change | 6 |
| IX. | Gases and Gas Laws | 7 |
| X. | Kinetic Molecular Theory | 2 |
| XI. | Atomic Theory | 6 |
| XII. | Bonding | 5 |
| XIII. | Solutions | 6 |
| XIV. | Equilibrium | 4 |
| XV. | Acids and Bases | 8 |
| XVI. | Oxidation-Reduction | 5 |
| XVII. | Electrochemistry | 4 |
| XVIII. | Nuclear Chemistry | 3 |
| XIX. | Environmental Chemistry (Optional) | |
| | Total: | 100 per[illegible] |

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*It is estimated that after subtracting for time consumed with [illegible] [illegible] tests, review sessions, laboratory work, etc., there are [illegible] 100 periods for classroom instruction in the academic year.

## CORRELATION OF ESSENTIAL GENERAL TOPICS WITH THEIR PRESENTATION IN THREE POPULAR TEXTBOOKS

High school chemistry texts now contain about 600 pages, so it can be a problem for teachers to decide what sections of the book to cover in depth and what material can be left out. The Education Committee has carefully examined the three most widely adopted texts in our state and have determined those chapters and pages that present the material that we believe to be truly essential. Suggested allocations of time to each topic are presented.

A teacher using one of these texts could simply follow the outline presented and be assured that he/she is presenting the most important material and allocating approximately the correct amount of time to each topic.

Teachers using other texts can easily construct similar sheets from our Essential General Topics list and a quick survey of the text being used.

TEXTBOOKS KEYED TO ESSENTIAL GENERAL TOPICS

Chemistry - A Modern Course
by: Smoot, Price, and Smith
Charles E. Merrill Publishing Company (1983)

| Topic | | Chapter | Page | Lecture Periods |
|---|---|---|---|---|
| I. | Introduction | 1 | 3-15 | 1 |
| II. | Metric System and Density | 2 | 17-23<br>34-35 | 6 |
| III. | Scientific Notation and Significant Figures | 2 | 23-29 | 4 |
| IV. | Naming, Symbols and Formulas | 4 | 58-68 | 7 |
| V. | Equations | 6 | 96-99 | 6 |
| VI. | Stoichiometry and the Mole Concept | 5<br>7 | 76-83<br>96-100 | 10 |
| VII. | Descriptive Chemistry | 10<br>11 | 174-194<br>200-218 | 10 |
| VIII. | Energy and Chemical Change | 7<br>*20 | 112-122<br>*397-407 | 6 |
| IX. | Gases and Gas Laws | 15<br>18<br>19 | 297-302<br>356-371<br>376-391 | 7 |
| X. | Kinetic Molecular Theory | 15<br>23<br>17 | 444-455<br>296-302<br>339-349 | 2 |
| XI. | Atomic Theory | 8<br>9 | 126-146<br>160-169 | 6 |
| XII. | Bonding | 12<br>*13<br>*14 | 224-239<br>248-262<br>274-286 | 5 |
| XIII. | Solutions | 5<br>21<br>22 | 82,83<br>412-423<br>426-434 | 6 |
| XIV. | Equilibrium | 23 | 444-461 | 4 |
| XV. | Acids and Bases | 24 | 468-481 | 8 |
| XVI. | Oxidation-Reduction | 26 | 514-529 | 5 |
| XVII. | Electrochemistry | 27 | 534-556 | 4 |
| XVIII. | Nuclear Chemistry | 28 | 562-581 | 3 |
| | | | Total Periods | 100 |

* NOTE - Optional Chapters (advanced topics)

TEXTBOOKS KEYED TO ESSENTIAL GENERAL TOPICS

Chemistry, The Study of Matter
by: Henry Dorin
Cebco Publishing Company (1982)

| Topic | Chapter | Page | Lecture Periods |
|---|---|---|---|
| I. Introduction | 1 | 2-8 | 1 |
| II. Metric System and Density | 2,3 | 10-32 | 6 |
| | | 134-136 | |
| III. Scientific Notation and Significant Figures | 2 | 18-22 | 4 |
| | 3 | 37-40 | |
| IV. Naming, Symbols and Formulas | 10 | 205-206 | 7 |
| | 13 | 250-264 | |
| | 23 | 450-451 | |
| V. Equations | 14 | 268-289 | 6 |
| VI. Stoichiometry and the Mole Concept | 6 | 112-113 | 10 |
| | 8 | 151-152 | |
| | 15 | 293-312 | |
| VII. Descriptive Chemistry | 12 | 212-245 | 10 |
| | 14 | 284-285 | |
| | 28 | 551-553 | |
| VIII. Energy and Chemical Change | 4 | 44-60 | 6 |
| | 7 | 117-127 | |
| | 14 | 276 | |
| | 20 | 382-285 | |
| IX. Gases and Gas Laws | 5 | 66-79 | 7 |
| | 6 | 81-112 | |
| X. Kinetic Molecular Theory | 6 | 107-116 | 2 |
| | 7 | 117-129 | |
| XI. Atomic Theory | 8 | 142-164 | 6 |
| | 9 | 175-184 | |
| XII. Bonding | 10 | 194-209 | 5 |
| | 11 | 212-224 | |
| XIII. Solutions | 16 | 320-329 | 6 |
| | 17 | 337-346 | |
| XIV. Equilibrium | 21 | 401-416 | 4 |
| XV. Acids and Bases | 23 | 434-458 | 8 |
| | 24 | 463-471 | |
| | 25 | 480-488 | |
| XVI. Oxidation-Reduction | 26 | 501-512 | 5 |
| XVII. Electrochemistry | 27 | 518-529 | 4 |
| | 28 | 532-545 | |
| XVIII. Nuclear Chemistry | 31 | 618-639 | 3 |
| | | Total Periods | 100 |

## TEXTBOOKS KEYED TO ESSENTIAL GENERAL TOPICS

Modern Chemistry
by: Metcalfe, Williams, and Castka
Holt, Rinehart and Winston (1982)

| Topic | Chapter | Page | Lecture Periods |
|---|---|---|---|
| I. Introduction | 1 | 1-7 | 1 |
| II. Metric System and Density | 1 | 7,8<br>13-16 | 6 |
| III. Scientific Notation and Significant Figures | 1 | 23-29 | 4 |
| IV. Naming, Symbols and Formulas | 2 | 41-43<br>45-46 | 7 |
| | 7 | 137-142 | |
| | 15 | 321-322<br>337-338 | |
| V. Equations | 2 | 50-55 | 6 |
| | 8 | 161-179 | |
| | 15 | 313-334 | |
| VI. Stoichiometry and the Mole Concept | 8 | 176-179 | 10 |
| VI. Descriptive Chemistry | 1 | 9-11 | 10 |
| | 5 | 89-107 | |
| | 8 | 168-175 | |
| | 15 | 335 | |
| | 23-30 | 523-626 | |
| VIII. Energy and Chemical Change | 1 | 11-13<br>18-21 | 6 |
| | 6 | 112-113<br>121-122 | |
| IX. Gases and Gas Laws | 10-11 | 203-241 | 7 |
| X. Kinetic Molecular Theory | 10 | 203-208 | 2 |
| | 12 | 244-245<br>252 | |
| XI. Atomic Theory | 3;4 | 629-648 | 6 |
| XII. Bonding | 6 | 109-133 | 5 |
| XIII. Solutions | 13 | 269-292 | 6 |
| XIV. Equilibrium | 21 | 461-489 | 4 |
| XV. Acids and Bases | 15 | 313-334 | 8 |
| XVI. Oxidation-Reduction | 22 | 493-520 | 5 |
| | 6 | 113-133 | |
| XVII. Electrochemistry | 22 | 507-520 | 5 |
| XVIII. Nuclear Chemistry | 31 | 629-648 | 3 |
| | | Total Periods | 100 |

## DEMONSTRATIONS AND EXPERIMENTS IN THIS BOOK KEYED TO ESSENTIAL GENERAL TOPICS

| | Topic | Demonstration Number | Experiment Number |
|---|---|---|---|
| I. | Introduction (Nature, Science, Chemistry) | 1 | 1 |
| II. | Metric System and Density | | 2 |
| III. | Scientific Notation and Significant Figures | 2 | 3 |
| IV. | Naming, Symbols, and Formulas | 3 | 4 |
| V. | Equations | 4 | 5 |
| VI. | Stoichiometry and the Mole Concept | 5,6 | 6 |
| VII. | Descriptive Chemistry | 7,8 | 7,8 |
| VIII. | Energy and Chemical Change | 9,10 | 9 |
| IX. | Gases and Gas Laws | 11 | 10 |
| X. | Kinetic Molecular Theory | 12 | |
| XI. | Atomic Theory | 13 | 11 |
| XII. | Bonding | | 12 |
| XIII. | Solutions | 14,15,16 | 13 |
| XIV. | Equilibrium | 17 | 14 |
| XV. | Acids and Bases | 18 | 6 |
| XVI. | Oxidation-Reduction | 19 | 15 |
| XVII. | Electrochemistry | | 16,17 |
| XVIII. | Nuclear Chemistry | | |
| XIX. | Environmental Chemistry (Optional) | 20 | |

See Chapter 7 page 1 for the titles and page numbers of the demonstrations.

See the beginning of Chapter 8 for the titles and page numbers of the experiments.

# LEARNING OBJECTIVES FOR HIGH SCHOOL CHEMISTRY KEYED TO ESSENTIAL GENERAL TOPICS

The objectives listed below are a minimum list, and are not intended to be a total list of learning objectives.

To assist persons writing test questions based on these objectives, it is intended that:

1. the term "define" means to give a definition of, identify the correct definition from a list of definitions, or complete a partial definition; or give a brief explanation.
2. the term "list" also means to select items from a list of items.
3. where applicable, the statement of an objective implies the reverse statement also. For example, if the objective calls for naming an element when given its symbol, it is implied that the symbol may be given and the name of the element requested.

An asterisk implies an advanced or optional topic.

## I. INTRODUCTION

A. Define chemistry and matter

B. List and/or arrange in order the steps in the scientific method.

C. Name five chemicals commonly used in the home.

D. Define the terms: element, atom, compound, and molecule.

E. Identify a material as either an element, a compound, a homogeneous mixture, or a heterogeneous mixture.

## II. METRIC SYSTEM AND DENSITY

A. Choose the most appropriate metric unit of mass, length, or volume of an object.

B. 1) Define the metric prefixes given in the set below by writing the prefix name, symbol, and equivalent.

   2) Differentiate between the metric prefixes by making conversions within each set of quantities:

      a) Mass -- kg, g, mg, μg

      b) Length -- km, m, cm, mm, μm, nm

      c) Volume -- L, mL, μL

      d) Energy -- cal, kcal or joule, kj

C. Identify the definition of mass, weight, volume, and density.

D. Identify the freezing point and boiling point of water on the Celsius scale.

E. Use the density formula to calculate density, mass, or volume when given the other two quantities.

F. Compare densities of two different substances.

G. Use the factor-label method (dimensional analysis) to obtain the desired unit in solving problems.

## III. SCIENTIFIC NOTATION AND SIGNIFICANT FIGURES

A. Express any number in scientific notation.

B. Make use of scientific notation in addition, subtraction, multiplication, and division.

C. Round off answers to the proper number of significant figures.

## IV. NAMING, SYMBOLS AND FORMULAS

A. Given the names of at least thirty common chemical elements, write the symbols.

B. Given certain monatomic and polyatomic ions, write the symbol and the charge for each.

C. Given the formulas, names, or ions of (a) common binary compounds and (b) compounds containing sulfate, sulfite, nitrate, nitrite, phosphate, carbonate, or ammonium ions write any of the others, i.e. formulas, names.

D. Write the names and formulas of common acids and bases.

## V. EQUATIONS

A. Write an equation in sentence form (word equation) when given a chemical equation.

B. Balance a simple chemical equation by inspection when given the formulas or names of all reactants and products.

C. Write and balance chemical equations for typical reactions, such as synthesis, decomposition, single displacement, and double displacement when given a set of reactants.

## VI. STOICHIOMETRY AND THE MOLE CONCEPT

A. Define "atomic mass"

B. Use the periodic table to determine the atomic mass of any element when given its symbol or the molecular mass of any element or compound when given its formula.

C. Calculate the number of moles in a given mass of a substance the formula of which is known.

D. Determine the number of atoms in a given mass of an element using Avogadro's number and the atomic mass.

E. Calculate the mass of an element when given Avogadro's number, the number of atom/s, and the atomic mass.

F. Given a fixed number of moles or the mass of a reactant or product and a balanced equation, calculate the mass or moles of either a product or a reactant.

G. Given the formula of a compound, calculate the percent of an element present in the compound.

*H. Given the percent composition of a compound, determine its empirical formula.

*I. Identify the limiting reagent in a mixture of reactants by calculating and comparing the number of moles of each reactant present using a balanced chemical equation.

## VII. DESCRIPTIVE CHEMISTRY

A. Identify an element as a metal, nonmetal, metalloid, or noble gas using a periodic table.

B. Identify various chemical families on the periodic table and their general properties.

C. Define the terms: element, atom, compound, and molecule.

D. Identify a material as either an element, a compound, a homogeneous mixture, or a heterogeneous mixture.

E. Given a list of properties, identify the properties as chemical or physical.

F. Given a list of changes, identify the changes as chemical or physical.

G. Identify general properties of gases, liquids and solids.

H. Given a list of solubility rules, predict if a precipitate is formed upon mixing solutions of known chemicals.

I. Sequence given elements in terms of relative activities using the periodic table.

J. Give a natural occurence or source of some common elements and compounds including oxygen, silicon, chlorine, ammonia, and sulfuric acid.

K. Using a periodic table, compare elements in terms of atomic radius, ionization energy, or electronegativity.

Learning Objectives for High School Chemistry
Keyed to Essential General Topics, page 4

## VIII. ENERGY AND CHEMICAL CHANGE

A. Define the terms: calorie and joule.

B. Use specific heats to calculate temperature change when a specific amount of energy is added to a given amount of a substance.

C. State the uses of a calorimeter.

D. Compute the energy required to change a given mass of a substance through melting, temperature change, and boiling (or vice versa), when given values for the heat of fusion, specific heat, and heat of vaporization.

E. Define endothermic and exothermic reactions.

F. Define endothermic and exothermic reactions in terms of bond energies.

*G. Determine heats of reaction using standard heats of formation.

## IX. GASES AND GAS LAWS

A. Define the absolute or Kelvin temperature scale.

B. Identify standard temperature in $^{o}$C and K, and standard pressure in atm, mm Hg, Pascal, and torr, recognizing that standard temperature and pressure are often referred to as STP or SC.

C. Identify the graphical relationships for gases between volume and temperature (Charles' law) and between volume and pressure (Boyle's law).

D. Solve problems using gas laws.

E. Define the molar volume of a gas.

F. Solve problems using the molar volume of a gas.

*G. Calculate the partial pressure of a gas in a mixture when given the total pressure and the pressure of all other gases present.

## X. KINETIC MOLECULAR THEORY

A. List the postulates of the kinetic-molecular theory.

B. Define pressure and diffusion of gases in terms of kinetic molecular theory.

C. Explain the effect of increasing intermolecular forces on density, melting point, or boiling point of a substance.

Learning Objectives for High School Chemistry
Keyed to Essential General Topics, page 5

## XI. ATOMIC THEORY

A. Identify various theories of the atom, including Rutherford, Bohr, and electron cloud theories by matching the theory to its description.

B. Identify the three fundamental particles of an atom when given the charge, mass, and location of the particle.

C. Determine the number of protons, electrons, or neutrons in an element when given the atomic number and the atomic mass of the element, or vice versa.

D. Write the electron configurations of any of the first 20 elements.

E. Identify the number of electrons in the outermost energy level of any A Group element.

F. Define isotopes.

G. Draw the electron-dot structure of any A Group element.

## XII. BONDING

A. Define ionic and covalent bonds.

B. Given a list of electronegativities and a pair of elements, identify the bond formed between the elements as ionic, polar covalent, or pure covalent.

C. Given a periodic table and a pair of elements, identify the bond formed between the elements as ionic, polar covalent, or pure covalent.

D. Given a list of properties, match the type of bonding with the correct properties.

E. Identify selected diatomic molecules as polar or nonpolar making use of the periodic table.

F. Draw and/or identify the structural formulas for $CH_4$, $H_2O$, $NH_3$, $CCl_4$, $CO_2$.

*G. Given simple molecules, identify or write their Lewis structures.

## XIII. SOLUTIONS

A. Express the concentration of a solution as percent by mass when given masses of the solute and solvent.

B. Given the percent by mass and the mass of the solution, determine the amount of each substance in the solution.

C. Calculate one of the following when given the other two: molarity, mass of solute, and volume of solution.

D. Match the following terms with definitions or examples: solute, solvent, electrolyte, nonelectrolyte, solubility, immiscible, miscible, saturated, supersaturated, unsaturated, molarity, polar, and nonpolar.

E. Given a pair of liquids, predict if they are miscible.

F. Given a combination of aqueous solutions, write the net ionic equation for the reaction.

G. List factors that affect solubility.

## XIV. EQUILIBRIUM

A. Define a reversible reaction.

B. Set up a mass action expression for an equilibrium reaction.

C. Using LeChatelier's principle, predict the effect upon a reaction at equilibrium of: (1) changing the temperature, (2) changing the concentrations of reactants and products, (3) changing the pressure, or (4) adding a catalyst.

D. Given a value of the equilibrium constant, determine if the forward or reverse reaction will be favored, or if there will be more reactants or products at equilibrium.

E. Given concentrations at equilibrium, calculate the equilibrium constant; or calculate an unknown concentration when given the equilibrium constant and concentration data.

## XV. ACIDS AND BASES

A. Given a list of aqueous solutions, identify those solutes which are strong electrolytes, weak electrolytes, or nonelectrolytes.

B. Identify the properties of acidic and basic solutions.

C. Given a list of statements describing theories of acids and bases, identify each theory as Arrhenius' or Bronsted-Lowry's.

D. Identify names or formulas for the strong acids; $HNO_3$, HCl, and $H_2SO_4$, and for the strong bases; NaOH and KOH.

E. Define the terms: strong acid, weak acid, strong base, weak base.

F. Given a list of acids and bases, identify the acids and bases as strong or weak.

G. Define pH.

H. Given the molar concentration of a strong acid or base, calculate the pH of the solution.

I. Write equations for neutralization reactions of strong acids and strong bases and name the salt formed.

J. Given titration data for reactions of strong acids and strong bases, determine unknown concentrations or volumes.

*K. Given a salt, predict if an aqueous solution of the salt would be acidic, basic, or neutral.

## XVI. OXIDATION-REDUCTION

A. Given the formula of a compound, identify the oxidation state of any element in it.

B. Given an oxidation-reduction reaction, identify the species oxidized, the species reduced, the oxidizing agent, or the reducing agent.

C. Given a list of chemical reactions, select those which are of the oxidation-reduction type.

*D. Given an oxidation-reduction reaction, balance the equation.

## XVII. ELECTROCHEMISTRY

A. Define a galvanic (voltaic) and an electrolytic cell.

B. Define: anion, cation, anode, cathode.

C. Given an electrochemical process, write half reactions occuring at the cathode and anode.

## XVIII. NUCLEAR CHEMISTRY

A. Identify $\alpha$, $\beta$, and $\gamma$ radiation with respect to mass and charge.

B. Given half-lives of isotopes, arrange in order of stability.

C. Complete and balance nuclear equations illustrating $\alpha$ or $\beta$ decay.

D. Define fusion and fission as related to the atomic nucleus.

*E. Describe the use of the concept of half-life in dating substances; i.e. carbon-14 dating.

*F. Explain the effect of radiation on biological systems.

## XIX. ENVIRONMENTAL CHEMISTRY

*A. List some probable causes and possible environmental consequences of acid rain.

*B. Define and discuss the greenhouse effect as related to the average temperature of the earth.

*C. List two ways by which the ozone layer may be depleted.

# SUGGESTED TOPICS FOR A SECOND-YEAR HIGH SCHOOL CHEMISTRY COURSE

by The Education Committee of the
Mississippi Section of the American Chemical Society

It is recommended that a second-year high school chemistry course consist of a study of special topics. Some suggested topics are listed below. Neither a year of organic chemistry nor a year of college-level general chemistry is advised for the second high school chemistry course.

## TOPICS

First Semester: Reinforce with a more sophisticated treatment important topics from first-year chemistry such as:

- Atomic Theory
- Bonding
- Energy and Chemical Change
- Ionic Equilibrium
- Kinetics
- Electrochemistry
- Solution Stoichiometry

Second Semester:

- Organic Chemistry (concluding with polymer chemistry)
  (Duration: Six Weeks)
- Qualitative Inorganic Analysis
  (Duration: Six Weeks)
- Environmental Chemistry
- A Chemistry Project

## CHEMISTRY FOR NON-SCIENCE-ORIENTED STUDENTS

Left unaddressed in this book, unfortunately, is the subject of what should be taught in high school chemistry courses for the benefit of those students who do not plan to go to college or who may not study chemistry even if they go to college. This very important subject should be a high priority concern of chemistry teachers everywhere. The most promising approach to the problem comes from the American Chemical Society (the national organization) which is developing an innovative, modular, interdisciplinary "chemistry and society" course entitled CHEMISTRY IN THE COMMUNITY (CHEMCOM). CHEMCOM is now being field tested and will soon be available for evaluation and possible use by high school chemistry teachers either as an alternative curriculum or to enrich more traditional courses. Each module will require about four to five weeks of teaching time. The CHEMCOM modules cover the following topics:

Supplying Our Water Needs
Feeding the World
Petroleum: To Burn or to Build?
Conserving Chemical Resources
The Chemical Industry: Promise and Challenge
Nuclear Chemistry in Our World
Chemicals, Air and Climate
Chemistry in Public and Personal Health

Much more information about CHEMCOM should be forthcoming within a year.

# CHAPTER 3

# TESTING STUDENTS

# CHAPTER 3 -- TESTING STUDENTS

## TESTS BASED ON TOPICS AND OBJECTIVES

If a teacher is to use learning objectives effectively, it is important that the tests given measure the students' learning of the material as set forth in the objectives. This does not mean, of course, simply rote memorization of facts. Questions can be worded in many different ways to elicit information to show that students have indeed accomplished a given learning objective.

## SAMPLE TESTS AND ANSWERS

In this chapter are found a number of "sample" tests that might be used when covering the material outlined in the Learning Objectives for High School Chemistry presented on pages 2-8 to 2-14. No attempt has been made to cover every objective within a given essential topic category, nor is every essential topic category covered. At the beginning of a group of sample questions, the number and title of the topic category from the Essential General Topics List on page 2-2, is given. To the left of each question is a letter corresponding to the letter of the learning objective under that topic category as found in the Objectives list. Answers to these sample questions are found immediately following the list of questions.

## I. INTRODUCTION

## TEST QUESTIONS

Objective

(B) 1. Arrange the following steps in the scientific method in the proper order:

hypothesize, make observations, draw conclusions, experiment, record data.

1. ______________________________
2. ______________________________
3. ______________________________
4. ______________________________
5. ______________________________

Answer questions 2-4 by putting the number of the answer in the blank.

(A) 2. Chemistry is the study of structure and _____ of substances.

(1) value (2) composition (3) energy (4) strength

(A) 3. All matter has ____.

(1) kinetic energy (2) radiant energy (3) mass (4) water

(A) 4. A chemist _____.

(1) Investigates the interaction between matter and energy.

(2) Investigates the composition and structure of matter.

(3) Explores space. (4) Does all of the above.

(C) 5. Fill in the blanks. Name at least five chemicals commonly used in the home.

(1) ________________________________________
(2) ________________________________________
(3) ________________________________________
(4) ________________________________________
(5) ________________________________________

## I. INTRODUCTION

## ANSWERS TO TEST QUESTIONS

1.1 make observation

1.2 hypothesize

1.3 experiment

1.4 record data

1.5 draw conclusions

2. (2)

3. (3)

4. (2)

5. answers vary

## II. METRIC SYSTEM AND DENSITY

## TEST QUESTIONS

Objective — Answer questions 1-12 by putting the number of the answer in the blank.

(C) 1. Which of the following properties of an object depends on the distance between the earth's crust and the object? _____

(1) mass (2) weight (3) density (4) volume

(Optional) 2. The SI unit of time is ______.

(1) minute (2) hour (3) second (4) year

(Optional) 3. The Kelvin scale is used to measure ______.

(1) pressure (2) mass (3) length (4) temperature

(D) 4. The number of degrees between the freezing point and boiling point of water on the Celsius scale is ______.

(1) 273° (2) 212° (3) 180° (4) 100°

(C) 5. Which is the term for mass per unit volume? ________

(1) area (2) density (3) speed (4) length

(F) 6. Comparing the densities of identical masses of cotton and lead, which statement is correct? _______

(1) 1 kg of cotton has a greater density.

(2) 1 kg of lead has a greater density.

(3) 1 kg of cotton and 1 kg of lead have equal densities.

(F) 7. Comparing the volumes of identical masses of cotton and lead, which statement is correct? ______

(1) 1 kg of cotton has the larger volume.

(2) 1 kg of lead has the larger volume.

(3) 1 kg of cotton and 1 kg of lead have equal volume.

Metric System and Density
Test Questions, page 2

Objective

(C) 8. The quantity of matter in an object has is called ________.

(1) inertia (2) momentum (3) mass (4) weight

(A) 9. Volume (the space occupied by an object) is measured in ____.

(1) liters (2) $dm^3$ (3) $in^3$ (4) all of these

(A) 10. The length (height) of a book is about ____.

(1) 25 m (2) 25 km (3) 25 cm (4) 25 mm

(A) 11. Which is the smallest mass? _____.

(1) pound (2) kilogram (3) milligram (4) nanogram

(F) (Optional) 12. Specific gravity of iron is 7.9, which means that any sample of iron has a mass 7.9 times as great as the mass of an equal volume of _____.

(1) air (2) water (3) aluminum (4) paper

(B) 13. Fill in the blanks:

| | Prefix | Symbol | Numerical Equivalent |
|---|---|---|---|
| (1) | micro | $\mu$ | ___ |
| (2) | _____ | k | $10^3$ |
| (3) | _____ | M | $10^6$ |
| (4) | nano | ___ | $10^{-9}$ |
| (5) | milli | m | ___ |
| (6) | _____ | c | $10^{-2}$ |

(B) 14. Make the following conversions

(1) 5 kg = _______ mg (2) 3 mm = _______ nm

(3) 1.5 L = _______ mL (4) 0.5 kcal = _______ cal

Metric System and Density
Test Questions, page 3

Objective

(G) 15. The answer to the following problem should be expressed in units of ________________.

$$\frac{52\ \text{weeks}}{\text{year}} \times \frac{7\ \text{days}}{\text{week}} \times \frac{24\ \text{hours}}{\text{day}} \times \frac{60\ \text{minutes}}{\text{hour}} \times \frac{60\ \text{seconds}}{\text{minute}} =$$

(E) 16. A piece of aluminum has a mass of 13.5 g and a volume of 5.0 mL. What is its density? __________

## II. METRIC SYSTEM AND DENSITY

## ANSWERS TO TEST QUESTIONS

1. (2)
2. (3)
3. (4)
4. (4)
5. (2)
6. (2)
7. (1)
8. (3)
9. (4)
10. (3)
11. (4)
12. (2)
13. (1) $10^{-6}$
    (2) kilo
    (3) mega
    (4) n
    (5) $10^{-3}$
    (6) centi
14. (1) $5 \times 10^{6}$ mg
    (2) $3 \times 10^{6}$ mn
    (3) $1.5 \times 10^{3}$ mL
    (4) $5 \times 10^{2}$ cal
15. second/year
16. 2.7 g/mL

## III. SCIENTIFIC NOTATION AND SIGNIFICANT FIGURES

### TEST QUESTIONS

Objective

(A) 1. Express the following numbers in scientific notation:

(1) 9270000 ____________

(2) 0.0052 ____________

(3) 485.29 ____________

(4) 0.000002 ____________

(5) 30,000 ____________

(A) 2. Express the following numbers in conventional form

(1) $7.05 \times 10^{-6}$ ____________

(2) $3.14 \times 10^{4}$ ____________

(3) $9.0 \times 10^{-5}$ ____________

(4) $-5.0 \times 10^{-2}$ ____________

(5) $6.002 \times 10^{-3}$ ____________

(B) 3. Multiply the following. Be sure to indicate the correct number of significant digits in your answer.

(1) 2.4 x 0.450 = ________

(2) 5000 x 0.150 = ________

(3) 0.007 x 0.05 = ________

(B) 4. Add the following indicating the correct number of decimal places.

(1)
30.0 g
0.542 g
4000.11 g

2)
0.0065 cm
0.502 cm
10 cm

Scientific Notation and Significant Figures
Test Questions, page 2

Objective

(B) 5. Perform the following operations:

(1) $(2.4 \times 10^{5}) \times (4.0 \times 10^{-3}) =$

(2) $\frac{1.4 \times 10^{-8}}{7.0 \times 10^{-3}} =$

(3) $(3.0 \times 10^{4}) + (6 \times 10^{3}) =$

(4) $(8.0 \times 10^{-4}) - (9.0 \times 10^{-6}) =$

(C) 6. How many significant digits should appear in the answers to the following problems?

(1) $4 \times .0056 =$

(2) $0.5 \times 8.75 \times 10^{2} =$

(3) $785 \times 65 =$

(4) $0.5 \div 0.05 =$

(5) $0.655 \div 9.56 \times 10^{1} =$

## III. SCIENTIFIC NOTATION AND SIGNIFICANT FIGURES

## ANSWERS TO TEST QUESTIONS

| | |
|---|---|
| 1.1 | $9.27 \times 10^{6}$ |
| 1.2 | $5.2 \times 10^{-3}$ |
| 1.3 | $4.8529 \times 10^{2}$ |
| 1.4 | $2 \times 10^{-6}$ |
| 1.5 | $3 \times 10^{4}$ |
| 2.1 | 0.00000705 |
| 2.2 | 31400 |
| 2.3 | 0.000090 |
| 2.4 | -0.05 |
| 2.5 | 0.006002 |
| 3.1 | 1.1 |
| 3.2 | 800 |
| 3.3 | 0.0004 |
| 4.1 | 4030.6 g |
| 4.2 | 11 cm |
| 5.1 | $9.6 \times 10^{2}$ |
| 5.2 | $2.0 \times 10^{-6}$ |
| 5.3 | $3.6 \times 10^{4}$ |
| 5.4 | $7.91 \times 10^{-4}$ |
| 6.1 | 1 |
| 6.2 | 1 |
| 6.3 | 2 |
| 6.4 | 1 |
| 6.5 | 3 |

## IV. NAMING, SYMBOLS AND FORMULAS

## TEST QUESTIONS

Objective Answer questions 1-7 by writing the number of the answer in the blank.

(C) 1. The correct formula for a substance composed of ions $Ag^+$ and $CO_3^{-2}$ is _____.

(1) $AgCO_3$ (2) $Ag_2CO_3$ (3) $Ag_3CO_3$ (4) $Ag_3(CO_3)_2$

(C) 2. _____is the formula of a binary compound.

(1) $HNO_3$ (2) $Al_2O_3$ (3) $NaOH$ (4) $O_2$

(C) 3. $CaCl_2$ is called ______.

(1) calcium chlorate (2) calcium chloride

(3) calcium chromate (4) calcium chlorite

(B) 4. If the algebraic addition of the charges on the atoms or ions in a compound add up to __________, the formula for the compound is correct.

(1) 4 (2) 2 (3) 0 (4) 1

(B) 5. An ion with an oxidation number of +2 can combine with two identical ions, each having an oxidation number of _____.

(1) +3 (2) +2 (3) +1 (4) -1

(C) 6. The Roman numeral II in the name copper(II) sulfate indicates _____.

(1) there are two copper ions per molecule

(2) there are two copper atoms per molecule

(3) there are two sulfate ions per molecule

(4) copper is in the +2 state

(C) 7. Which of the following formulas is incorrect? _____

(1) $NaS$ (2) $AlPO_4$ (3) $Ca(OH)_2$ (4) $KNO_3$

Naming, Symbols and Formulas
Test Questions, page 2

Objective

(D) 8. Write the common names next to the formula of the following acids.

(1) $HCl$ ____________________ (2) $H_2SO_4$ ____________________

(3) $HNO_3$ ____________________ (4) $H_2CO_3$ ____________________

(B) 9. Write the symbol and charge for each of the following ions.

(1) potassium ____________ (2) mercury(II) ____________

(3) hydrogen sulfate ____ (4) magnesium ____________

(5) phosphate ____________ (6) aluminum ____________

(C) 10. Name the following compounds:

(1) $Fe(OH)_2$ ____________________ (2) $SbCl_5$ ____________________

(3) $Cr_2O_3$ ____________________ (4) $N_2O_4$ ____________________

(5) $CCl_4$ ____________________ (6) AgF ____________________

(A) 11. Write the name of the element next to its symbol.

(1) Fe ______________ (6) He ______________

(2) Cl ______________ (7) Bi ______________

(3) I ______________ (8) Al ______________

(4) Hg ______________ (9) Mn ______________

(5) P ______________ (10) Ca ______________

## IV. NAMING, SYMBOLS AND FORMULAS

## ANSWERS TO TEST QUESTIONS

1. (2)
2. (2)
3. (2)
4. (3)
5. (4)
6. (4)
7. (1)
8. (1) hydrochloric acid
    (2) sulfuric acid
    (3) nitric acid
    (4) carbonic acid
9. (1) $K^+$
    (2) $Hg^{+2}$
    (3) $HSO_4^-$
    (4) $Mg^{+2}$
    (5) $PO_4^{-3}$
    (6) $Al^{+3}$
10. (1) iron(II) hydroxide
    (2) antimony(V) chloride
    (3) chromium(III) oxide
    (4) nitrogen(IV) oxide
    (5) carbon(IV) chloride
    (6) silver fluoride
11. (1) iron
    (2) chlorine
    (3) iodine
    (4) mercury
    (5) phosphorus
    (6) helium
    (7) bismuth
    (8) aluminum
    (9) manganese
    (10) calcium

## IX. GASES AND GAS LAWS

### TEST QUESTIONS

**Objective** — Answer questions 1-9 by putting the number of the answer in the blank.

(A) 1. A temperature of -20°C is equivalent to ____.

(1) -253 K (2) 80 K (3) 253 K (4) 293 K

(B) 2. A pressure of 380 mm of Hg is equivalent to _____.

(1) 0.5 atm (2) 1.0 atm (3) 2.0 atm

(4) none of the above

(A) 3. What is the relationship between degrees Celsius and Kelvin units?

(1) Celsius degrees are larger than Kelvin units.

(2) Celsius degrees are smaller than Kelvin units.

(3) Celsius degrees are the same as Kelvin units.

(A) 4. The temperature at which all molecular motion very nearly ceases is _____.

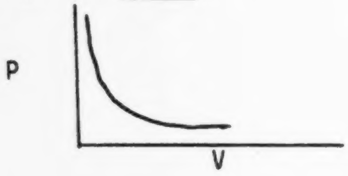
(1) 0°C (2) 100°C (3) 0 K (4) 100 K

(C) 5. The graph to the right shows that the relationship between the pressure and the volume of a gas is _____.

(1) direct

(2) indirect

(3) inverse



(C) 6. Charles' Law states that if a gas is subjected to constant pressure, the temperature varies directly with the _____.

(1) mass (2) density (3) number of molecules (4) volume

Gases and Gas Laws
Test Questions, page 2

Objective

(C) 7. Under which conditions of temperature and pressure will a sample of gas have the largest volume? _____

(1) 100 K and 300 mm Hg (2) 100 K and 150 mm Hg

(3) 200 K and 150 mm Hg (4) 200 K and 300 mm Hg

(F) 8. The volume of 2.0 mole of $O_2$ gas at STP is _____.

(1) 2.0 L (2) 11.2 L (3) 22.4 L (4) 44.8 L

(G) 9. A sample of gas was collected over water at standard pressure. If the water vapor pressure was 25 mm Hg, the pressure of the dry oxygen was _____.

(1) 25 mm Hg (2) 735 mm Hg (3) 760 mm Hg (4) 785 mm Hg

Complete the four statements below with one of these responses:

(1) increases (2) decreases (3) remains the same

(C) 10. A weather balloon is released. As the balloon rises, its size _________________ due to decreased atmospheric pressure.

(C) 11. As an automobile tire warms up when a car is driven, the pressure of the air in the tire _____________ because of the increase in temperature.

(C) 12. If a gas is enclosed in a rigid container and the container is cooled, the pressure exerted by the gas _______________.

(C) 13. When a gas is allowed to expand from 1 liter to 5 liters the number of moles of gas ________________ .

Gases and Gas Laws
Test Questions, page 3

Objective — Questions 14-17 involve calculations.

(D) 14. A gas collected when the pressure is 800 mm Hg has a volume of 700 mL. What volume, in milliliters, will the gas occupy at 500 mm of pressure if the temperature remain constant?

(D) 15. A gas measures 400 mL at a temperature of 10°C. What volume will the gas occupy at 20°C if the pressure remains constant?

(D) 16. A gas collected at a temperature of 27°C and a pressure of 80 mm Hg occupies a volume of 500 mL. Calculate the volume of the gas in milliliters at 17°C and 700 mm pressure.

(D) 17. How many moles of any gas at 2.00 atm. and 35°C are contained in an 8.00 liter container?

## IX. GASES AND GAS LAWS

## ANSWERS TO TEST QUESTIONS

1. (3)
2. (1)
3. (3)
4. (3)
5. (2)
6. (4)
7. (3)
8. (4)
9. (2)
10. increases
11. increases
12. decreases
13. remains the same
14. 1120 mL
15. 414 mL
16. 552 mL
17. 0.63 mole

## XI. ATOMIC THEORY

## TEST QUESTIONS

Answer questions 1-9 by putting the number of the answer in the blank.

Objective

(A) 1. Which of the following statements is not included in Dalton's Atomic Theory? _____

(1) All matter is composed of atoms.

(2) Atoms cannot be separated into parts.

(3) An element can have more than one kind of atom.

(4) Atoms of different elements differ from each other.

(A) 2. The scientist who discovered the electron while experimenting with gas discharge tubes was _____.

(1) Mullikan (2) Rutherford (3) J. J. Thomson (4) Dalton

(C) 3. The charge on the nucleus of a sodium atom is _____.

(1) +1 (2) -11 (3) +11 (4) +23

(C) 4. What atom is used as the standard for the atomic mass scale? _____

(1) hydrogen-1 (2) carbon-12 (3) oxygen-16 (4) nitrogen-17

(B) 5. Which particles are contained in the nucleus of an atom? _____

(1) protons and electrons (2) electrons and neutrons

(3) protons and neutrons (4) protons, neutrons, and electrons

(A) 6. A unique set of wavelengths absorbed or emitted by a particular substance is _____.

(1) a spectrum (2) the frequency (3) an x-ray

(4) the photoelectric effect

Atomic Theory
Test Questions, page 2

Objective

(F) 7. Isotopes of an element are atoms which differ in _____.

(1) the number of protons (2) the atomic number

(3) the number of electrons (4) the number of neutrons

(A) 8. An atomic orbital is _____.

(1) a negatively charged particle (2) a wavelength

(3) a circular path around the nucleus

(4) a highly probable location in which an electron may be found in an atom

(D,E) 9. The maximum number of electrons in the third energy level of an atom is _____.

(1) 2 (2) 8 (3) 18 (4) 32

(C) 10. Fill in the blanks.

| Name of Nuclide | Atomic Number | Mass Number | Composition of Nucleus | | Number of Electrons |
|---|---|---|---|---|---|
| | | | Protons | Neutrons | |
| lithium-7 | 3 | | | | |
| oxygen-17 | | | 8 | | |
| aluminum-27 | | | | | 13 |
| phosphorus-30 | 15 | | | | |
| phosphorus-32 | | | | | |

Atomic Theory
Test Questions, page 3

**Objective**

(D,G) 11. Fill in the blanks.

| Chemical Symbol | Atomic Number | Electron Configuration Notation | Electron Dot Notation |
|---|---|---|---|
| B | 5 | | |
| F | 9 | | |
| Mg | 12 | | |
| Si | 14 | | |
| Ar | 18 | | |

Questions 12-14 are essay questions.

(A) 12. How does the size of the nucleus of an atom compare with the size of the atom?

(F) 13. Using the Bohr picture of an atom, draw and describe the structure of each of the three hydrogen isotopes.

(B) 14. (a) What are the main parts of an atom?

(b) What particles are found in each part?

(c) Describe each type of particle.

## XI. ATOMIC THEORY

## ANSWERS TO TEST QUESTIONS

1. (3)
2. (3)
3. (3)
4. (2)
5. (3)
6. (1)
7. (4)
8. (4)
9. (3)
10.

| Name of Nuclide | Atomic Number | Mass Number | Composition of Nucleus | | Number of Electrons |
|---|---|---|---|---|---|
| | | | Protons | Neutrons | |
| Lithium-7 | 3 | 7 | 3 | 4 | 3 |
| Oxygen-17 | 8 | 17 | 8 | 9 | 8 |
| Aluminum-27 | 13 | 27 | 13 | 14 | 13 |
| Phosphorus-30 | 15 | 30 | 15 | 15 | 15 |
| Phosphorus-32 | 15 | 32 | 15 | 17 | 15 |

11.

| Chemical Symbol | Atomic Number | Electron Configuration Notation | Electron Dot Notation |
|---|---|---|---|
| B | 5 | $1s^2 2s^2 2p^1$ | $\cdot\mathrm{B}:$ |
| F | 9 | $1s^2 2s^2 2p^5$ | $:\overset{..}{\underset{.}{\mathrm{F}}}:$ |
| Mg | 12 | $1s^2 2s^2 2p^6 3s^2$ | $\mathrm{Mg}:$ |
| Si | 14 | $1s^2 2s^2 2p^6 3s^2 3p^2$ | $:\mathrm{Si}:$ |
| Ar | 18 | $1s^2 2s^2 2p^6 3s^2 3p^6$ | $:\overset{..}{\underset{..}{\mathrm{Ar}}}:$ |

Atomic Theory
Key to Test, page 2

12. The nucleus of an atom is one ten thousandth ($10^{-4}$) the size of the atom.

| | | | |
|---|---|---|---|
| 13. | 1P | 1P | 1P |
| | | 1N | 2N |
| | 1 | 2 | 3 |

H-1 or protium has 1 proton in the nucleus and one electron moving about it.

H-2 or deuterium has 1 proton and 1 neutron in the nucleus and 1 electron moving about it.

H-3 or tritium has 1 proton and 2 neutrons in the nucleus and 1 electron moving about it.

14. (a) The nucleus and shells or energy levels.

(b) Proton(s) and neutron(s) are found in the nucleus. Electron(s) are found in the shell(s) or energy level(s)

(c) Protons are positively charged particles with one unit of positive charge and have a mass of $1.673 \times 10^{-24}$ g. (or 1.0 atomic mass units)

Neutrons are neutral particles and have a mass of $1.675 \times 10^{-24}$ g. (or 1.0 atomic mass units)

Electrons are negatively charged particles with one unit of negative charge and have a mass of $9.110 \times 10^{-24}$ g. (or approximately 0.0 atomic mass units)

## CONSTRUCTION AND USE OF QUESTION BANKS

It is suggested that, once a chemistry teacher has decided on a definite list of objectives, he/she begin to develop a data bank of test questions for each item on the objectives list. For example, under topic category IV (Naming, Symbols and Formulas), item B, one would develop a number of questions designed to test the students' knowledge of that item. Then a teacher has only to reach into the question bank under that topic and item and pull out any question. For another test or examination over the same material, another question would be used. The question bank could be updated and expanded annually and used year after year. A simple approach to developing a question bank is to write a question on an index card, write the correct answer on the back, and maintain a file box of questions appropriately indexed so that one can find immediately questions dealing with a particular item in a particular topic category. If the teacher has access to a personal computer, a faster, more sophisticated (although not necessarily better) system can be developed. Teachers are urged to swap question bank material and even to purchase commercially available question banks to be weeded and personalized for use. No teacher should have to spend hours generating tests; neither should a teacher dash off a test of superficial, trivial questions. A good question bank is the answer!

## A METHOD FOR RAPIDLY PREPARING TESTS FROM A QUESTION BANK

by Dr. Howard P. Williams
Director of General Chemistry
University of Southern Mississippi

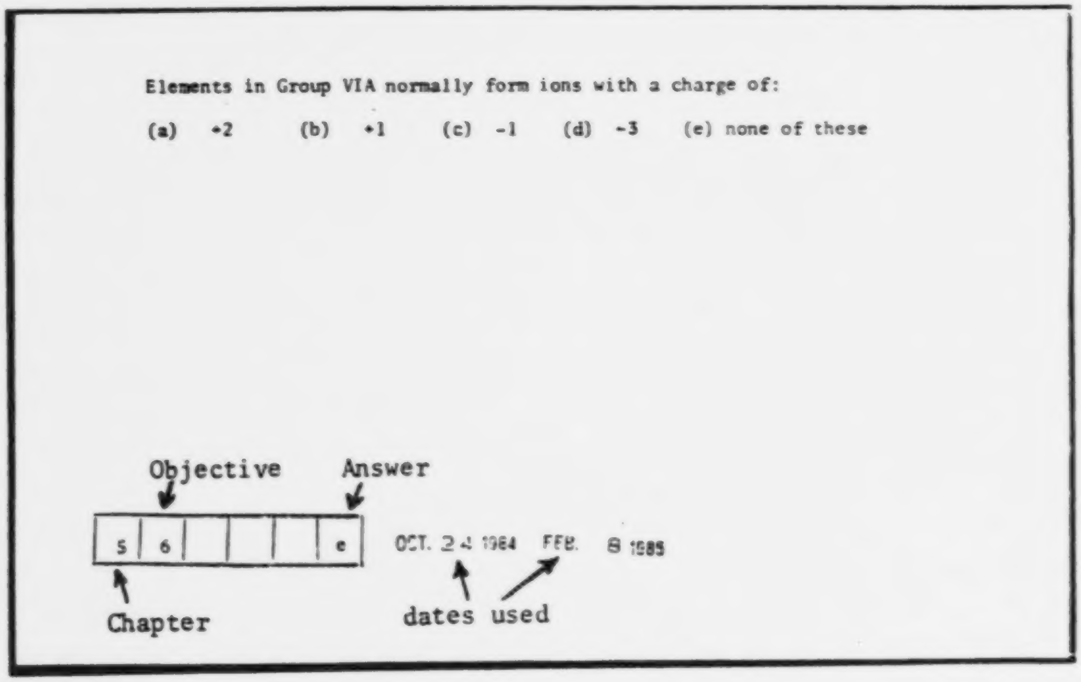
This method, while extremely simple, requires the availability of a copying machine. The question bank consists of a generous supply of 5" x 8" index cards and a box in which to file them. The format of a typical card is shown in figure 1. The first line is the typed or hand-printed question with a left-hand margin of one inch. At the bottom of the card is a block of information about the question (i.e. learning objective number, chapter, difficulty, etc.) along with the correct answer. The use of tabbed index cards affords a convenient and organized way of keeping the question topics separated in the file box.

Clear acetate folders or 8 1/2" by 11" mylar page protection covers may be used to assemble a test master using the index cards. The clear plastic folder serves to hold the overlapping cards as shown in figure 2. A sheet of typing paper behind the cards prevents the black marks that might result with only the 8" wide cards being copied. The cards are arranged in the plastic folder with the left edge of the 5" x 8" card inserted in the fold of the folder. Then the card with the second question is lapped over the bottom of the first card with enough space between the bottom of the first card and the top of the second card to allow the student to write his/her answer. The card with the third question is lapped over the bottom of the card with the second question, and so forth until a full page of questions is assembled.

The clear plastic folder is then closed and the enclosed assembly of overlapping cards is placed question side down on the copying machine. The copy results in a complete page of questions but without question numbers which can then be typed or written to the left of each question. This page is the master copy of the test and copies for the students are made from it. If the test is more than one page long, the additional pages are generated in the same manner.

This procedure permits the rearrangement of the same questions for several forms of the same examination with very little time expended. It is especially useful for furnishing "example tests" for students to use in studying; however, when first building the basic bank of test questions, it is unlikely that the teacher would have enough diverse questions to furnish practice tests.

In order to prevent unwanted duplication of the same questions in successive tests, the date is written at the bottom of the card each time the question is used. The back of each card can be used to show how the question is answered and any mathematical steps needed to solve a problem. Thus by simply turning the cards over and repeating the process described above, an answer sheet can be easily generated from the same cards. For multiple choice questions, the information furnished on the front of the cards is usually all that is needed.



Elements in Group VIA normally form ions with a charge of:

(a) +2 (b) +1 (c) -1 (d) -3 (e) none of these

| 5 | 6 | | | | e |
|---|---|---|---|---|---|

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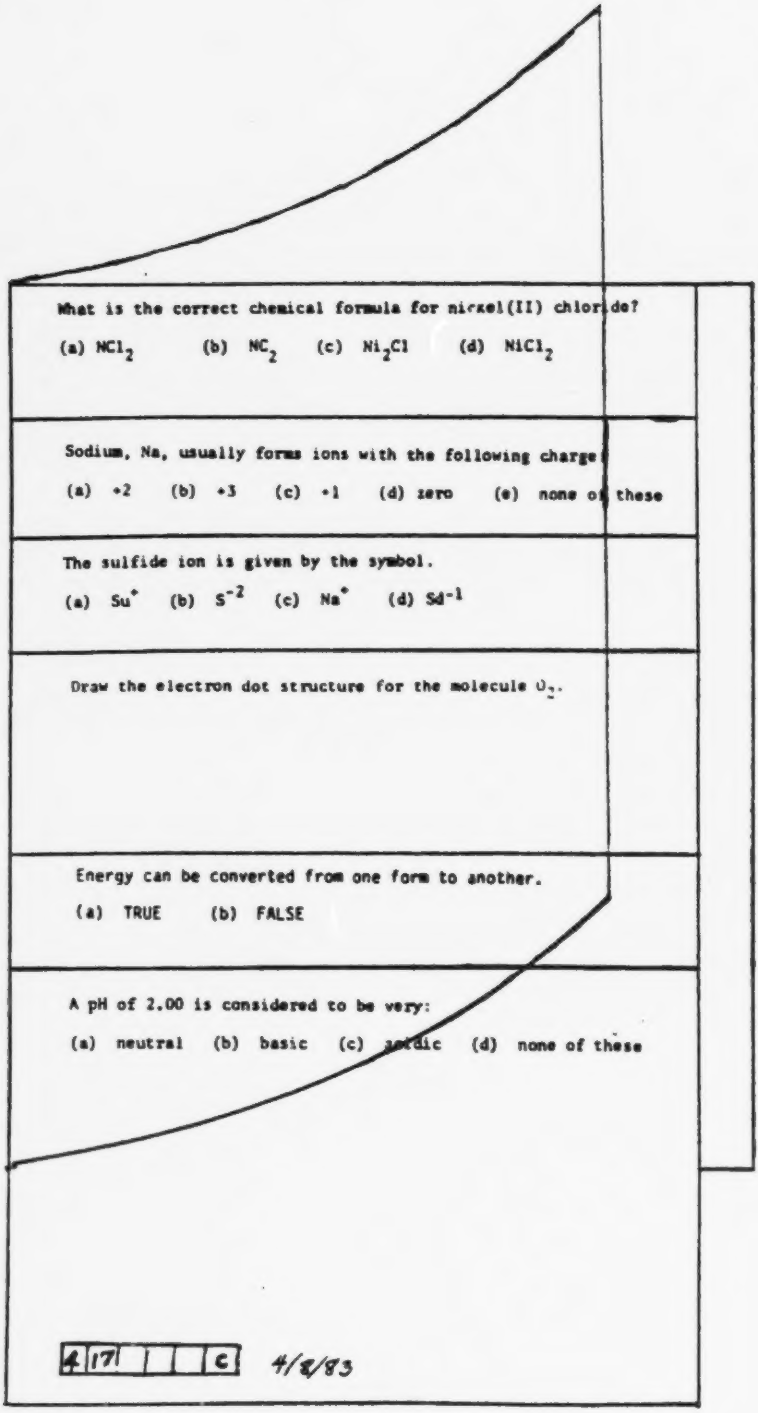
Objective
Answer
Chapter
dates used

Figure 1



What is the correct chemical formula for nickel(II) chloride?

(a) $NCl_2$ (b) $NC_2$ (c) $Ni_2Cl$ (d) $NiCl_2$

Sodium, Na, usually forms ions with the following charge:

(a) +2 (b) +3 (c) +1 (d) zero (e) none of these

The sulfide ion is given by the symbol.

(a) $Su^+$ (b) $S^{-2}$ (c) $Na^+$ (d) $Sd^{-1}$

Draw the electron dot structure for the molecule $O_2$.

Energy can be converted from one form to another.

(a) TRUE (b) FALSE

A pH of 2.00 is considered to be very:

(a) neutral (b) basic (c) acidic (d) none of these

| | | | | | |
|---|---|---|---|---|---|
| 4 | 17 | | | | c |

4/8/83

Figure 2

# CHAPTER 4

# THE LABORATORY

# CHAPTER 4 -- THE LABORATORY

## INTRODUCTION

The laboratory is where chemistry comes to life. It is where most practicing chemists first came to like the subject. It offers one of the few opportunities students have to learn to make observations, organize information logically, draw conclusions and present results. A chemistry course without laboratory experiences for the students is at best an incomplete, inferior course. A teacher does his/her students a great disservice if no effort is made to include laboratory learning

The reasons for not offering a laboratory are numerous and most are valid. Laboratory preparation and cleanup does involve extra work, a laboratory is expensive, accidents can happen in a laboratory, often laboratory equipment is poor. To which we respond: but laboratory instruction is so important to the learning of chemistry, it must be done.

In a manner analogous to the development of the learning objectives presented earlier, the Education Committee has compiled a list of Minimum Skills for the Chemistry Laboratory. These skills should be mastered by everyone who completes a course in high school chemistry. The list is so short and the skills so simple that a good teacher will want to make this list only the beginning of a more comprehensive list of laboratory skills.

Following the list of laboratory skills, two sample tests on laboratory learning are included. One is a "practical" test; the other is a written test. Questions on both tests are keyed to the list of Minimum Skills for the Chemistry Laboratory.

Next there are some suggestions concerning student laboratory notebooks and reports followed by detailed procedures for preparing reagent solutions.

The chapter concludes with some suggestions for estimating the cost of incorporating new experiments into the laboratory program and detailed price lists of chemicals and supplies for all the experiments described in this handbook.

Price lists for the demonstrations described in this handbook have not been prepared because only small quantities of supplies are involved and most items are readily available.

## MINIMUM SKILLS FOR THE CHEMISTRY LABORATORY

1. Identify the common pieces of laboratory equipment and be able to spell the names correctly.
2. Demonstrate a knowledge of laboratory safety rules.
3. Point out the location and operation of fire extinguishers and other safety devices.
4. Demonstrate the proper use of a balance.
5. Demonstrate the proper use of a Bunsen burner.
6. Measure the mass of less than 200 g of a substance to within 0.01 g.
7. Measure the length of an object to within 0.1 cm.
8. Measure the volume of about 10 mL of a liquid to within 0.1 mL using a graduated cylinder.
9. Demonstrate the proper method for removing a small amount of solid from a bottle.
10. Demonstrate the proper method for pouring a small amount of a corrosive liquid from a bottle.
11. Measure the temperature of a liquid using a Celsius thermometer.
12. Measure the volume of a liquid delivered by a buret and/or pipet.
13. Determine the barometric pressure by using a barometer (if available).
14. Demonstrate how to separate a mixture of a solid and a liquid by filtration.
15. Make proper observations and write an understandable description of laboratory activities.
16. Organize information in a logical manner.
17. Derive conclusions based on observations using data or graphs.

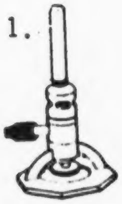
## SUGGESTED QUESTIONS FOR A WRITTEN LABORATORY TEST

(Keyed to the List of Minimum Skills for the Chemistry Laboratory)

Skill Number

(1)
1. ____________________
2. ____________________
3. ____________________
4. ____________________
5. ____________________
6. ____________________
7. ____________________
8. ____________________
9. ____________________
10. ____________________
11. ____________________
12. ____________________
13. ____________________
14. ____________________
15. ____________________

1.



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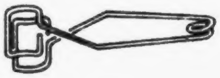

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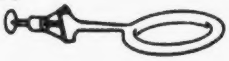

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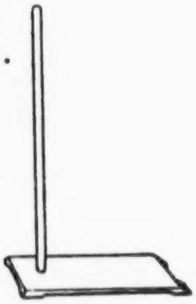

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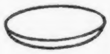

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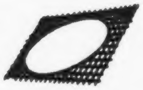

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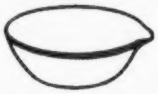

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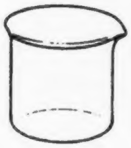

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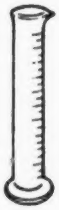

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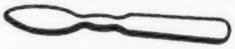

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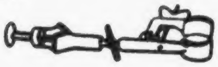

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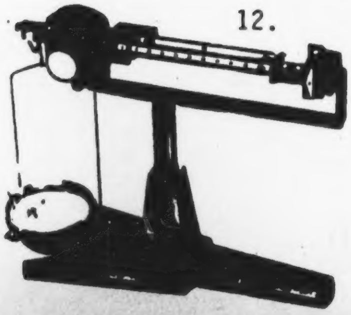

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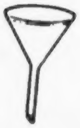

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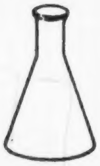

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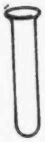


Suggested Questions for a
Written Laboratory Test, page 2

Skill Number

(3) 16. What are the locations of the fire extinguishers and first aid kit?

(2) 17. Name one piece of safety equipment that is worn during every laboratory period.

(2) 18. What should be done immediately if you spill acid or other chemicals on your skin?

(1,7) 19. If you were told to obtain 10 mL of hydrochloric acid, what equipment would you use?

(4,5) 20. Describe how you would weigh:

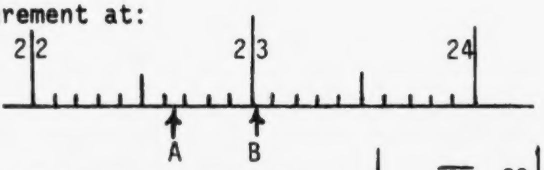
(a) 11.5 g of salt

(b) 12 g of water

(6) 21. What is the measurement at:

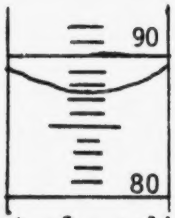
(a) arrow A?

(b) arrow B?



(4) 22. What type (color) of bunsen burner flame is best for laboratory work and why?

(7) 23. What is the volume of the liquid in the graduated cylinder to the left?



(8) 24. Describe how you would remove a small amount of a solid chemical from a reagent bottle?

(11) 25. A student did an experiment to determine the percent of water in a potato. She weighed an evaporating dish and found the mass to be 60.75 g. She placed some pieces of potato in the dish and found the mass of dish and potato to be 110.90 g. After heating the dish and potato until all the water was removed the mass of dish and potato was 70.78 g. Set up a data table for this data and calculate and record the mass of and percent of water in the potato.

(12) 26. Draw a graph to represent the following data.

| Mass | Volume |
|---|---|
| 1 g | 5 $cm^3$ |
| 3 g | 15 $cm^3$ |
| 5 g | 25 $cm^3$ |
| 7 g | 35 $cm^3$ |

## ANSWERS TO SUGGESTED QUESTIONS FOR A WRITTEN LABORATORY TEST

1. Bunsen burner
2. Test tube holder
3. Ring clamp
4. Ring stand/ring
5. Watch glass
6. Wire gauze
7. Evaporating dish
8. Beaker
9. Graduated cylinder
10. Spatula
11. Test tube clamp/utility clamp/buret clamp
12. Balance
13. Funnel
14. Erlenmeyer flask
15. Test tube
16. Description of locations
17. Safety glasses/safety goggles
18. Flush with an ample supply of water. Contact the teacher.
19. Graduated cylinder
20. (a) Make sure the pointer on the balance is set at zero. Place a piece of weighing (glazed) paper on the balance pan and determine its mass. Record the mass of the paper. Set the balance to read the mass of the paper plus 11.5 g. Using a spatula, obtain a quantity of salt from a bottle and carefully pour the salt onto the weighing paper until the pointer of the balance comes to zero.

    (b) Repeat procedure above except use a beaker instead of weighing paper.

Answers to Suggested Questions
for a Written laboratory Test, page 2

21. (a) 22.65

(b) 23.02

22. Blue, because it burns hotter and cleaner.

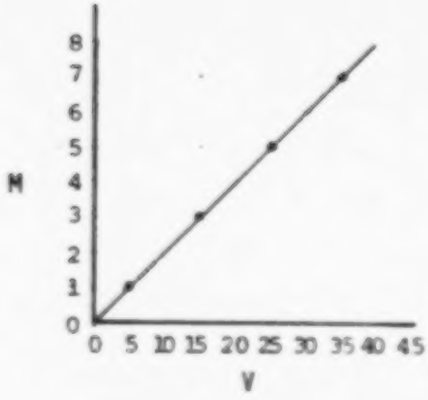
23. 87.5 mL.

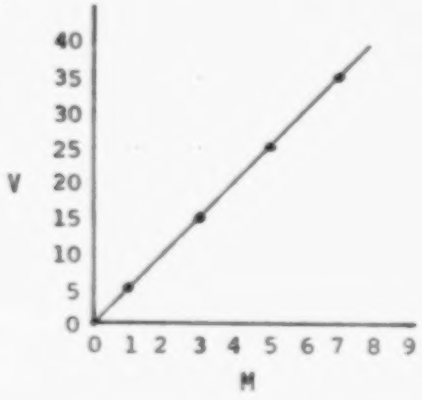
24. First read the label on the bottle to make sure you are getting the chemical requested. Using a spatula remove the desired amount of salt and put it on a glazed paper. Do not put any surplus salt back in the reagent bottle; instead discard it.

25.

| | | |
|---|---|---|
| Mass of dish + fresh potato | | = 110.90 g |
| Mass of dish | | = 60.75 g |
| Mass of fresh potato | 110.90 g - 60.75 g | = 50.15 g |
| Mass of dish + dry potato | | = 70.78 g |
| Mass of dry potato | 70.78 g - 60.75 g | = 10.03 g |
| Mass of water | 50.15 g - 10.03 g | = 40.12 g |
| % of water in potato (40.12 - 50.15 g) x 100 | | = 80.00% |

26.





## SUGGESTED QUESTIONS FOR A PRACTICAL LABORATORY TEST

(Keyed to the List of Minimum Skills for the Chemistry Laboratory)

Skill Number

(1) 1. Items of equipment have been placed on laboratory tables with a number beside each piece of equipment. Identify each item of equipment by writing its name by the appropriate number.

(1) ____________________
(2) ____________________
(3) ____________________
(4) ____________________
(5) ____________________
(6) ____________________
(7) ____________________
(8) ____________________
(9) ____________________
(10) ____________________
(11) ____________________
(12) ____________________
(13) ____________________
(14) ____________________

(3) 2. Describe the position of the fire extinguisher and first aid kit in this room.

(4,6) 3. Weigh 7.60 g of sodium chloride using the balance.

(5) 4. Light and adjust a Bunsen burner to produce a flame appropriate for laboratory use.

(7) 5. Obtain a meter stick and determine the length (height) of your chemistry textbook in centimeters.

(8) 6. Use a graduated cylinder to measure 15.6 mL of tap water.

(9) 7. Use a spatula to remove approximately 3 g of sodium chloride from the reagent bottle.

(10) 8. Pour 15.0 mL of hydrochloric acid into a 50.0 mL graduated cylinder (or 5.0 mL into a 10.0 mL graduated cylinder).

Note to the Teacher: There can be no accompanying answer sheet for this test which requires observation of student skills.

# REPORTING LABORATORY RESULTS

There are many ways to structure the presentation of laboratory results. Individual reports may be required for each experiment, fill-in-the-blank report sheets could be provided, or each student could keep a laboratory notebook. No matter which choice is made, the students should be informed of the required format before the first experiment.

A laboratory notebook should be hardbound (not spiral or looseleaf) with ruled paper. When graphs are required, commercial graph paper should be used. The graphs may be taped or stapled into the notebook. All entries in the notebook should be neat, orderly, and legible. However, experimental data should not be recopied. If a change is necessary, a line should be drawn neatly through the data to be changed, and the new data entered.

Shown below is a suggested format for the student to use in recording information about an experiment in the laboratory notebook. Alternatively, this format could be used for individual reports submitted by the student to the teacher. Note that the report of each experiment is divided into three sections.

Pre-laboratory:

1. The date of the experiment
2. The title of the experiment
3. A statement of the objective of the experiment
4. The pre-laboratory assignments
5. An outline of the experimental procedure

Laboratory:

1. A detailed record of all observations
2. A record of all measurements made in the experiment
3. Names and quantities of materials used

Post-Laboratory:

1. Answers to all questions asked about the experiment
2. Each calculation shown in detail
3. A statement that summarizes the success or failure of the experiment. This can include error calculations.
4. Information and answers relating to "further study" assignments

# REAGENT SOLUTION PREPARATION

Reagent solutions must be prepared properly if laboratory work is to be successful. Procedures for preparing reagent solutions are relatively simple but require practice for proficiency. As a review, presented below are detailed instructions for the preparation of two common laboratory solutions; 0.10 M $CuSO_4 \cdot 5H_2O$ and 1.0 M HCl. The techniques described are general and can be used to prepare various other solutions.

I. For the preparation of 1.0 L of 0.10 M $CuSO_4 \cdot 5H_2O$, the following procedure is suggested:

1. Obtain the needed chemicals, glassware, and equipment:
    - (a) $CuSO_4 \cdot 5H_2O$ (solid)
    - (b) Distilled water
    - (c) Balance, 0.01 g sensitivity
    - (d) Beaker, 250 ml
    - (e) Volumetric flask, 1.0 L, with stopper
    - (f) Funnel
    - (g) Stirring rod
    - (h) Spatula
    - (i) Plastic wash bottle

2. Clean glassware with detergent and tap water, rinse well with tap water, and rinse finally with 2 or 3 portions (20-15 mL each) of distilled water.

3. Calculate the mass of $CuSO_4 \cdot 5H_2O$ needed to prepare 1.0 L of 0.10 M $CuSO_4 \cdot 5H_2O$ (molar mass or gram formula weight = 249.68).

$$\text{g } CuSO_4 \cdot 5H_2O = 1.0 \text{ L} \times \frac{0.10 \text{ mol } CuSO_4 \cdot 5H_2O}{\text{L}} \times \frac{249.68 \text{ g } CuSO_4 \cdot 5H_2O}{\text{mol } CuSO_4 \cdot 5H_2O}$$

$$= 24.97 \text{ g } CuSO_4 \cdot 5H_2O$$

4. Weigh out 24.97 g of $CuSO_4 \cdot 5H_2O$ into a 250 ml beaker.

5. Add 100-150 mL distilled water to the beaker and stir with a stirring rod to dissolve the $CuSO_4 \cdot 5H_2O$.

6. Pour the $CuSO_4 \cdot 5H_2O$ solution through a funnel into a 1.0 L volumetric flask.

Care should be taken to transfer all the $CuSO_4 \cdot 5H_2O$ solution from the beaker to the flask by rinsing the beaker with several portions of distilled water or by washing the side of the beaker carefully with distilled water from a plastic wash bottle, adding the wash water to the volumetric flask also. Finally, rinse the funnel several times before removing it.

7. Add distilled water to the volumetric flask to bring the liquid level up to near the mark on the neck of the flask.
8. Mix the contents of the flask by swirling and then allow the liquid to reach room temperature.
9. Use a plastic wash bottle to bring the level of solution exactly to the mark. Cap the flask with the stopper provided.
10. Mix the contents of the flask by inverting the flask and swirling. Repeat the inverting and swirling process at least 10 times.
11. Transfer the solution to a clean glass or plastic bottle for storage.
12. Label the bottle "0.10 M $CuSO_4 \cdot 5H_2O$". Put the date of preparation on the label. Initial the label.

II. A procedure for the preparation of 250 mL of 1.0 M HCl:

1. Obtain the needed chemicals, glassware, and equipment.
    - (a) Concentrated hydrochloric acid, 12 M
    - (b) Distilled water
    - (c) Beaker, 100 mL
    - (d) Graduated cylinder, 10 mL
    - (e) Volumetric flask, 250 mL
    - (f) Funnel on ring stand
    - (g) Plastic wash bottle
2. Clean all glassware with detergent and tap water, rinse well with tap water, and rinse finally with 2 or 3 portions (10-15 mL each) of distilled water.

3. Calculate the volume of concentrated (12M) HCl needed to prepare 250 mL of 1.0 M HCl. Use this formula:

$$V_1 \times M_1 = V_2 \times M_2$$

$V_1$ = volume of 12 M HCl needed

$M_1$ = 12 M

$V_2$ = 250 mL

$M_2$= 1.0 M

$$V_1 = \frac{V_2 \times M_2}{M_1} = \frac{250 \text{ mL} \times 1.0 \text{ M}}{12 \text{ M}} = 20.8 \text{ mL}$$

4. Using a funnel, carefully pour slightly more than 21 mL of concentrated HCl into the 250 mL beaker.
5. Transfer the concentrated HCl from the beaker to a 25 mL graduated cylinder to give a volume of exactly 20.8 mL. Discard the excess concentrated HCl by rinsing down the drain with a large quantity of water.
6. Add approximately 100 mL of distilled water to a 250 mL volumetric flask.
7. Pour the 20.8 mL of concentrated HCl through a funnel into the 250 mL volumetric flask. Care should be taken to transfer all the HCl from the graduated cylinder to the flask by rinsing the cylinder with 2 or 3 small portions of distilled water, and also rinsing the funnel before removing it.
8. Add distilled water to the flask to bring the liquid level up to near the mark on the neck of the flask.
9. Mix the contents of the flask by swirling and then allow the contents to reach room temperature.
10. Use a plastic wash bottle to bring the level of solution exactly to the mark. Cap the flask with the stopper provided.
11. Mix the contents of the flask by inverting the flask and swirling. Repeat the inverting and swirling process at least 10 times.
12. Transfer the solution to a clean glass or plastic bottle for storage.
13. Label the bottle "1.0 M HCl". Put the date of preparation on the label. Initial the label.

## ESTIMATING THE COST OF LABORATORY INSTRUCTION

Before incorporating a new experiment or demonstration into the chemistry course, a teacher will want to know the approximate cost of the chemicals, supplies, and equipment involved.

Usually it will be possible to reduce the cost of supplies by using common sense and making some obvious substitutions. For example, if the procedure for an experiment you are thinking of using reads "Pour 250 mL of water into a 400 mL beaker," it would be foolish to buy 400 mL beakers if your laboratory is already stocked with 500 mL beakers and if the slightly larger beaker would serve the purpose of the experiment just as well. Sometimes a satisfactory substitution may not be as obvious. For example, an experiment might require the use of an accurate timer, which costs about $70. However, among any group of high school students today, it is almost certain that several will be wearing digital wristwatches with a stopwatch function. These watches are probably more accurate than the timer and their use costs nothing!

To help the teacher who is considering using some of the experiments described in this handbook, a list of the approximate cost of chemicals and supplies for each experiment is presented on the following pages. The prices are 1985 retail prices for individual items. Buying in quantity will reduce the costs significantly. "State contract" prices, if applicable to purchases by your school, will further reduce costs.

It is suggested that chemicals and supplies be bought for only a few new experiments each year (but on a regular year-to-year basis) rather than a huge one-time purchase of items for a large number of new experiments. This approach will allow you to make a more careful evaluation of the experiments in the hands of your students. Moreover, in the future, additional new experiments can be incorporated to make use of many of the items of chemicals and supplies already on hand for the few initial new experiments you adopted.

# CHEMICALS AND SUPPLIES LISTS

## EXPERIMENT 1. THE BURNING OF A CANDLE

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| lampblack | 500 g | $14.00 | many |
| sugar, cube | 1 lb. | .60 | 100 |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| beaker, 400 mL | 1.81 |
| beaker, 1000 mL | 4.34 |
| burner, Bunsen | 7.20 |
| candle | .27 |
| cardboard | free |
| cotton ball | .01 |
| flask, Erlenmeyer, 500 mL | 3.00 |
| forceps | .65 |
| wire gauze | 1.00 |
| watch glass | .95 |
| match | .15 |
| paper, plain, white | free from school office |
| watch or timer | 70.00* |

* a wristwatch with stopwatch function may be used

## EXPERIMENT 2. MASS AND VOLUME RELATIONSHIP

### A. CHEMICALS

| | |
|---|---|
| aluminum metal, 13 mm dia x 5.1 cm long specific gravity specimen | $1.25 each |
| brass metal, 13 mm dia x 5.1 cm long specific gravity specimen | 1.25 each |
| unknown metal | 1.25 each |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| graduated cylinder, 100 mL | 6.80 |
| balance, 200 g, 0.01 g sensitivity | 112.00* |
| thread | .20 |

*One balance per 4 stations

## EXPERIMENT 3. SCIENTIFIC NOTATION AND INDIRECT MEASUREMENTS

### A. CHEMICALS

NONE

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| balance, 200 g, 0.01 g sensitivity | 112.00* |
| 50 or 100 mL beaker | 1.60 or 1.80 |
| evaporating dish, small | 3.94 |
| medicine dropper | .15 |
| ream of paper | (borrow from school office) |
| ruler, metric | .80 |

* One balance per 4 stations

## EXPERIMENT 4. SYMBOLS AND FORMULAS

### A. CHEMICALS

NONE

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| cardboard or heavy paper | $ .75 |
| graph paper | 1.00 |
| scissors | 3.00 |

## EXPERIMENT 5. CHEMICAL EQUATIONS

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quanity Purchased |
|---|---|---|---|
| barium chloride | 100 g | 5.30 | many (>200) |
| hydrochloric acid | 2.37 L | $ 12.95 | many (>200) |
| magnesium sulfate | 500 g | 19.75 | many (>500) |
| sugar | 5 pounds | 1.60 | many (>1500) |
| magnesium ribbon | 12 | 9.95 | (for teacher's demonstration) |
| 1984 or later penny | 1 ea | .01 | .5 |

(continued on next page)

Chemicals and Supplies List, page 3

Experiment 5, continued

## B. SUPPLIES (PER STATION)

| | | |
|---|---|---|
| burner, Bunsen | $ 7.20 | |
| graduated cylinder, 10 mL | 4.90 | |
| graduated cylinder, 100 mL | 6.80 | |
| file, triangular | 3.50 | |
| wire gauze | .95 | |
| watch glass | .95 | (for teacher's demonstration) |
| test tube holder | 2.50 | |
| litmus paper, red | .50 | (for teacher's demonstration) |
| ring with clamp | 4.00 | |
| ring stand | 13.00 | |
| test tube, 25 mm x 200 mm | .73 | |
| test tube, 16 x 150 mm | .29 | |
| test tube holder | 2.50 | |
| tongs | 6.80 | (for teacher's demonstration) |

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## EXPERIMENT 6. EXPERIMENTAL STOICHIOMETRY

## A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| ethyl alcohol | 500 mL | 3.20 | many (>800) |
| hydrochloric acid | 2.37 L | 12.95 | many (>100) |
| phenolphthalein | 25 g | 8.35 | 100 |
| sodium hydroxide | 500 g | 10.35 | 40 |
| sulfuric acid | 4 kg | 25.50 | many (>100) |
| distilled water | 1 gal | .65 | 30 |

## B. SUPPLIES (PER STATION)

| | | |
|---|---|---|
| beaker, 50 mL | 1.60 | |
| graduated cylinder, 5 or 10 mL | 4.90 | |
| medicine dropper | .15 | |
| six 20x150 mm test tubes | 2.64 | (.44 each) |

Chemicals and Supplies List, page 4

## EXPERIMENT 7. INVESTIGATION OF HARD AND SOFT WATER

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Statopms Served By Minimum Quantity Purchased |
|---|---|---|---|
| calcium chloride | 100 g | $ 5.95 | 14 |
| calcium sulfate | 100 g | 9.90 | 40 |
| distilled water | 1 gal | .65 | 200 |
| liquid detergent | 80 oz. | .89 | 100 |
| magnesium chloride | 100 g | 7.60 | 14 |
| sodium carbonate | 100 g | 13.85 | 200 |
| solid soap | 1 bar | .60 | 50 |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| 1 balance, 200 g cap.,0.01 g sens. | 112.00* |
| 3 test tubes. 13 x 100 mm | $ .69 |
| 1 test tube rack | 5.15 |

*One balance per 4 stations

---

## EXPERIMENT 8. IONS AND THEIR REACTIONS

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| barium chloride | 100 g | $ 5.30 | 190 |
| copper(II) sulfate pentahydrate | 100 g | 5.90 | 163** |
| hydrochloric acid | 2.37 L | 12.95 | 11,376 |
| iron(III) chloride | 100 g | 5.60 | 246 |
| lead(II) nitrate | 100 g | 4.75 | 120 |
| silver nitrate | 113 g | 76.00 | 265 |
| sodium carbonate | 500 g | 13.85 | 446 x 5 = 2,230 |
| sodium hydroxide | 500 g | 10.35 | 5,000 |
| sodium phosphate | 500 g | 4.90 | 1,220 |
| distilled water | 1 gal | .65 | 17 |

(continued on next page)

Chemicals and Supplies List, page 5

Experiment 8, continued

B. SUPPLIES (PER STATION)

| | |
|---|---|
| One spot plate | $ 6.63 |
| Nine small bottles with screw caps | 16.58 |
| Nine medicine droppers | 1.35 |

*This is calculated value. Any waste will decrease number of stations served.

$$** \#\text{Stations} = \frac{1 \text{ station}}{2.5 \text{ ml solu.}} \times \frac{10^3 \text{ ml}}{1 \text{ mole}} \times \frac{\text{mole}}{2.44.6 \text{ g}} \times 100 \text{ g} \sim 163$$

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## EXPERIMENT 9. ENERGY OF A CHEMICAL CHANGE

A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| sodium hydroxide | 2.5 Kg | $23.10 | 400 |
| sulfuric acid | 4 Kg | 25.50 | many (>1400) |

B. SUPPLIES (PER STATION)

| | |
|---|---|
| balance, 200 g cap 0.01 g sensitivy | 112.00* |
| styrofoam cup | .05 |
| graduated cylinder, 100 mL | 6.80 |
| flask, Erlenmeyer, 250 mL | 2.74 |
| thermometer, -20 to 100°C, 1°div. | 6.90 |

*One balance per 4 stations

---

## EXPERIMENT 10. THE MOLECULAR MASS OF A GAS

A. CHEMICALS

NONE

B. SUPPLIES PER STATION

| | |
|---|---|
| balance | 112.00* |
| butane lighter | $ .60 |
| graduated cylinder, 100 mL | 6.80 |
| flask, Erlenmeyer, 250 mL | 2.74 |
| marking pencil | .65 |
| glass plate or watch glass | .95 |
| thermometer 20°C - 100°C, 1°div. | 6.90 |
| trough, metal | 7.50 |

*One balance per 4 stations

Chemicals and Supplies List, page 6

## EXPERIMENT 11. ISOTOPIC MASS VS. ATOMIC MASS

### A. CHEMICALS

NONE

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| navy beans | $ 0.90/lb |
| pinto beans | 0.90/lb |
| blackeyed peas | 0.90/lb |
| 3 400 mL beakers | 5.43/4 beakers |
| 1 200 g cap. balance with 0.01 g sensitivity | 112.00* |

* One balance per 4 stations

---

## EXPERIMENT 12. ELECTRICAL CONDUCTIVITY AND CHEMICAL BONDING

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|---|
| calcium chloride | 100 g | 5.95 | | 100 |
| ethyl alcohol | 500 mL | 5.95 | | 5 |
| lead acetate | 100 g | 8.20 | | 100 |
| potassium bromide | 100 g | 5.95 | | 200 |
| sodium chloride | 1 lb | .27 | (from grocery) | 453 |
| sugar | 1 lb | 1.29 | (from grocery) | 453 |
| distilled water | 1 gal. | .65 | (from grocery) | 37 |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| 9-volt battery | .90 |
| conductivity cell components | |
| 2 paper clips | free from school office |
| cork | .04 |
| 4 dram vial | .20 |

## EXPERIMENT 13. PREPARATION OF A STANDARD SOLUTION

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| copper sulfate | 100 g | $ 5.90 | 33 |
| distilled water | 1 gal | .65 | 75 |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| balance 200 g cap, 0.01 g sensitivity | 112.00* |
| two beakers, 100 mL | 3.32 (1.66 each) |
| wash bottle | 1.20 |
| graduated cylinder, 25 mL | 5.40 |
| medicine dropper | .15 |
| sheet of white paper | free from school office |
| stirring rod | .23 |
| 2 test tubes, 13 x 100 mm | .46 (.23 each) |

*One balance per 4 stations

---

## EXPERIMENT 14. APPLYING "STRESS" TO A SYSTEM IN EQUILIBRIUM

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served Bu Minimum Quantity Purchased |
|---|---|---|---|
| hydrochloric acid | 2.37 L | 12.95 | many (>500) |
| ice | 5 lbs | 1.00* | 25 - 30 |
| potassium nitrate | 125 g | $ 12.25 | 60 |
| sodium chloride | 2000 g | 17.95* | 100 |
| distilled water | 1 gal | .65 | many (>500) |

*school cafeteria may supply this free of charge

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| beaker, 250 mL | $ 1.52 |
| graduated cylinder, 10 mL | 4.90 |
| medicine dropper | .15 |
| stirring rod | .23 |
| 2 test tubes, 13 x 100 mm | .46 (.23 each) |
| thermometer, -20 to 100°C, 1°C | 5.90 |

## EXPERIMENT 15. OXIDATION-REDUCTION REACTIONS OF METALS AND METAL IONS

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| copper metal (wire) | 4 oz | 4.85 | many |
| copper(II) nitrate (cupric nitrate) | 25 g | 5.40 | about 10 |
| iron metal | 100 g | 9.75 | about 40 |
| iron(II) sulfate (ferrous sulfate) | 500 g | 6.10 | about 40 |
| lead metal | 500 g | 8.85 | about 40 |
| lead(II) nitrate | 100 g | 4.75 | about 40 |
| magnesium metal (ribbon) | | | many |
| magnesium nitrate | 100 g | 5.60 | about 40 |
| zinc metal | 500 g | 7.65 | many |
| zinc nitrate | 100 g | 5.15 | about 40 |
| distilled water | 1 gal | .65 | 37 |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| paper for labels | free from school office |
| 9 test tubes (15mm x 125mm) | 2.79 (.31 each) |
| test tube rack | 5.65 |

---

## EXPERIMENT 16. ELECTROLYSIS OF POTASSIUM IODIDE IN WATER

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| potassium iodide | 100 g | $ 7.70 | 300 |
| (or sodium iodide) | (25 g) | 6.45 | (75) |
| copper wire, #13-18 | 10 meters | 9.75 | 50 |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| 2 battery, lantern 1.5 V D.C. | 8.40 |
| (9 volt battery) | (.90) |
| buret stand with clamp | 13.00 + 6.50 |
| litmus paper, blue | .30 |
| litmus paper, red | .30 |
| U-tube | 3.25 |

Chemicals and Supplies List, page 9

## EXPERIMENT 17. ELECTROPLATING COPPER

### A. CHEMICALS

| Item | Minimum Quantity Purchased | Cost | Maximum Number Of Stations Served By Minimum Quantity Purchased |
|---|---|---|---|
| copper(II) sulfate | 100 g | $ 5.90 | 20 |
| copper wire | 230 ft | 10.95 | 28 |
| ethanol | 4 L | 9.85 | many (>700) |
| sulfuric acid | 4 Kg | 25.50 | many (>2700) |

### B. SUPPLIES (PER STATION)

| | |
|---|---|
| batteries, 1.5 V D.C. Lantern | 8.40 |
| beaker, 250 mL | 1.52 |
| paper clip | free from school office |
| sand paper, fine | .50 |
| balance, analytical (optional) | |

# CHAPTER 5

# CHEMICAL SAFETY

# CHAPTER 5 -- CHEMICAL SAFETY

## INTRODUCTION

There is now in the mid-1980s an increased demand nationwide for better science teaching in the high schools with the expectation of more laboratory work by the students. Ironically this demand for more and better laboratory experiences by the students comes at the same time that the American public is being sensitized to the hazards of chemicals. Many teachers as well as students actually have a fear of working with chemicals in the laboratory. To compound the problem, we are living in what may be termed the age of the lawsuit. A chemistry teacher faces a very real possibility of being sued if a student is injured in the laboratory.

Thus the conscientious teacher who wants to increase and improve the laboratory experiences of the students must face the facts that many chemicals are dangerous, that accidents can happen, and that he/she could conceivably face a lawsuit if a student is injured. What is a teacher to do?

There is no alternative -- the teacher must conduct a chemistry laboratory. He/she must know the chemistry involved well enough not only to feel comfortable in working with the chemicals but confident in the teaching of students in the laboratory. Everything possible must be done to conduct a safe laboratory operation and to prevent accidents. Few situations are as horrible for a teacher to contemplate as that of a child seriously injured and possibly disfigured for life as a direct result of the teacher's action or lack of action. Ethically and morally it is the obligation of the teacher to do everything possible to make the laboratory a safe place for students to learn.

Before the students begin any experiment, the teacher should develop an awareness of any possible hazards in the experiment. It should not be assumed that an experiment described in a textbook (especially one more than a few years old) is safe. Each experiment should be carefully evaluated and approved by the teacher prior to incorporation into the laboratory program. Even if there appears to be very little risk to the students' safety, the teacher is morally and legally bound to inform the students, prior to the beginning of an experiment, of possible hazards and the precautions to be taken.

This handbook offers a number of relatively simple experiments that pose little risk. Potential hazards are clearly marked. A teacher's guide that accompanies each experiment frequently gives additional safety information. Yet, regardless of how safe the individual experiments are, if they are performed in an unsafe environment by students unaware of safety rules, the potential for serious trouble exists. This chapter on safety provides information about creating a good laboratory environment in which experiments can be conducted safely.

## GENERAL GUIDELINES FOR THE TEACHER IN THE CHEMISTRY LABORATORY

The teacher:

1. is expected to protect the health, welfare, and safety of his students.
2. is expected to foresee reasonable consequences of his actions or inactions.
3. must carefully plan and preview all activities.
4. should create an environment in which appropriate laboratory behavior is maintained.
5. should keep adequate records covering all areas of laboratory operation.
6. should never leave the laboratory when the students are present.
7. should convey information about all potential hazards to administrative personnel and insist that corrective measures be taken.
8. should be aware of local laws and regulations that relate to laboratory activities in chemistry.
9. must carefully instruct the students in laboratory safety techniques giving careful, clear directions before allowing the students to attempt any experiment.
10. must be careful to inform the students of any possible risk in performing a given laboratory procedure before they begin their experiment.

Note that the last two guidelines directly involve the students in the responsibility for their own safety. When the teacher has adequately instructed the students regarding hazards, the teacher's liability has been minimized.

## STUDENT INVOLVEMENT IN SAFETY PROGRAMS

The foundation for all safety activity should be a formal program having the official approval of the school's administrators and the active support of parents. Policies and guidelines may come from the administration, but the teacher has to assume the responsibility for making the school's safety program work; i.e. instructional objectives for safety should comprise the framework of all activities, with the teacher serving as the role model.

The best laid plans for safety or any other activity will never be realized if they do not involve the people for whom the program is designed. Teachers and student representatives should work in concert with the administration in the development of safety guidelines.

Teachers and students are usually very creative and resourceful in planning and conducting campaigns of various types. A very practical and worthwhile project for science classes is the production of a safety manual to which each student can make a contribution according to his/her capabilities, interests and needs. Such a manual might include a description of equipment and directions for using it. It should also include current information about occupational safety and health, sources of information, etc. Another useful group activity is the maintenance of a "chemistry bulletin board". Students would supply items for posting.

## LABORATORY AIDES

Since teachers will find it very difficult to do all of the work necessary to assure safety in the school laboratory, he/she should train reliable students to assist with routine activities and to make the teacher aware of special safety problems. Arrangements should be made by teachers and administrators for the orientation of student aides who will be involved in setting up experiments, making solutions, or keeping the chemistry stockroom in order.

Ideally an aide should have completed first year chemistry. It would be even better if the laboratory aide has taken second-year chemistry.

Of course, there are some precautionary measures that should be observed.

A student aide should never:

1. work alone.
2. work with dangerous chemicals.
3. work with large quantities of chemicals.
4. be left in charge of a chemistry laboratory class.

A student aide should be properly instructed and continually monitored. He/she must always wear the appropriate safety protection for the activity being pursued.

## FACILITIES AND EQUIPMENT REQUIRED FOR A SAFE LABORATORY

A high school chemistry laboratory should meet several minimal facility and equipment requirements in order to provide a safe environment for teaching laboratory chemistry. These minimal requirements are basic to all laboratory activities, but certainly do not address all the safety needs of many specific experiments. Proper facilities and safety equipment do not preclude accidents; however, they may limit their severity.

The items listed below are considered to be basic facility requirements for safety in the chemistry laboratory:

fire extinguisher

eyewash fountain

emergency shower

first aid kit

A fire extinguisher, preferably an ABC dry-type extinguisher, should be located near an exit. Ideally another extinguisher should be located near the back of the laboratory. Both should be well marked.

An eyewash fountain should be readily accessible and capable of delivering a constant water flow for at least fifteen minutes. Flinn Scientific, Inc. sells a kit for converting a laboratory faucet to an eyewash fountain. The kit costs less than $75.00, requires no new plumbing and can be installed by the teacher. A short hose attached to a faucet and equipped with an aerator device on the end is a minimally adequate substitute for an eyewash fountain. Squeeze-bottle eyewash stations containing various agents are not recommended as substitutes.

An emergency shower can be of several designs provided it is easily accessed and operated by a disoriented, injured person. The shower should deliver copious amounts of water in a wide dispersion in order to rapidly extinguish fires or dilute chemicals.

The first aid kit should be permanently mounted in a highly visible, easily reached location.

Physical facilities should meet the following criteria:

* Adequate ventilation including at least one fume hood which is needed for many common laboratory experiments involving flammable or toxic substances.

* Two exits

* Master control valve for the gas supply to laboratory stations.

* Reasonable access to the electrical breaker/fuse box that supplies the laboratory.

## PERSONAL PROTECTION

The human eye is most vulnerable to injury; therefore, every effort should be made to protect the eyes of anyone in a chemistry laboratory. Mississippi state law requires every student, teacher, and visitor(s) to wear eye protection at all times while participating in or observing activities in the chemistry laboratory.

Prescription glasses are not acceptable as eye protection. Safety goggles should be worn over prescription glasses. Contact lenses should never be worn in the laboratory. The student who normally uses contact lenses should be required to purchase a pair of prescription glasses for use (under safety goggles) while in the laboratory. If contact lenses are absolutely medically necessary and correctional prescription glasses cannot be substituted for them (a rare situation), it is recommended that a physician's statement of this fact be required from the student/parents and that the parents be advised of the increased risk the student assumes. Any student who wears contact lenses in the laboratory should be known to the teacher who will have special problems in coping with an eye injury to this person should an accident occur.

Below is a list of some other guidelines for laboratory safety:

1. Gloves appropriate to the activity should be worn as needed. When handling chemicals, use only gloves that will resist penetration by the chemical being handled.
2. Laboratory aprons or coats should be worn.
3. Shorts are prohibited.
4. Footwear should cover feet completely; no open-toe shoes or sandals.
5. Loose clothing (e.g. sleeves, full cut blouses, neckties, etc.), long hair, and dangling jewelry should be avoided; and, if worn, restrained with clips, rubber bands, etc.
6. Never work alone in a science laboratory or storage area.
7. Never eat, drink, smoke, chew gum or tobacco in the laboratory. Do not store food or beverages in the laboratory environment.
8. Wash hands after work in the science laboratory, and after spill cleanups.
9. Use safety shields or screens whenever there is a potential danger that an implosion or explosion of apparatus might occur.

10. Do not taste chemicals unless directed by the teacher to do so.

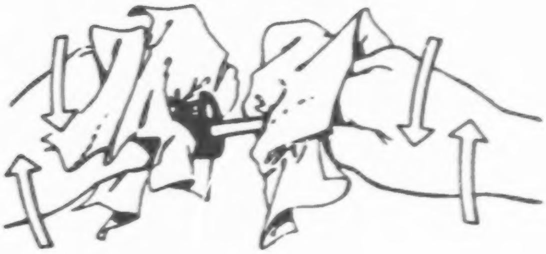
11. Do not touch chemicals unless directed by the teacher to do so.

These guidelines are not all-inclusive. They are merely some of the most obvious suggestions for personal protection in the chemistry laboratory.

## PERFORMING SOME BASIC LABORATORY OPERATIONS SAFELY

### A. Inserting Tubing, Thermometer, or Glass Funnel Into a Rubber Stopper

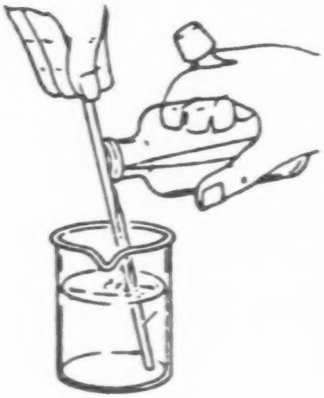
Be certain the ends of the glass are fire polished. Begin by lubricating the tip of the glass with glycerol or mineral oil. Never force the tubing into the stopper. Ease it in with a gentle twisting motion. Protect your hands with a cloth or a leather glove. If the insertion is difficult, place a cork borer that just slips over the glass into the cork hole. Slide the glass into position and extract the cork borer.



## B. Handling Liquid Reagents

Many reagent bottles do not have flat-top glass stoppers. In using such containers, remove the glass stopper and hold it between the fingers of the hand used to grasp the reagent bottle. Never lay a non-flat-top stopper from a reagent bottle on a laboratory bench because it will become contaminated and transfer the contamination to the contents of the bottle.

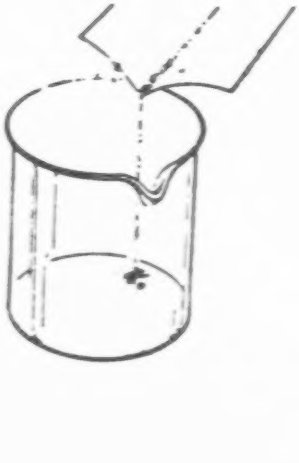
## C. Transferring Chemicals From Reagent Bottles



Remove the stopper as explained above in B. To transfer a liquid from one container to another hold a stirring rod against the lip of the vessel containing the liquid and pour down the rod which should be touching the inside of the receiving vessel.

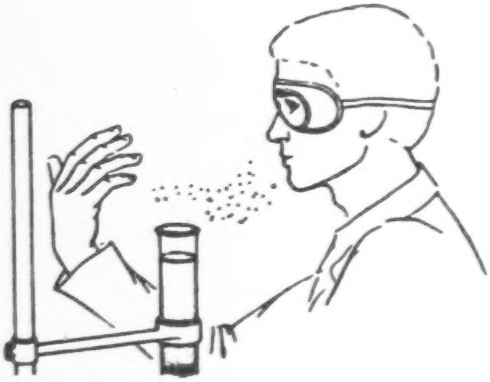
## D. Transferring a Solid

If the container has a screw cap, remove it and set it on the bench top, inner side facing up. Hold the bottle with the label against your hand, tilt the bottle, and roll it back and forth. Dispense the chemical onto waxed (glazed) paper. Try not to dispense more reagent than needed. If the solid is caked and does not pour easily, use a clean spatula with a long handle to break up the solid. Thoroughly clean and dry the spatula before using it in another container.



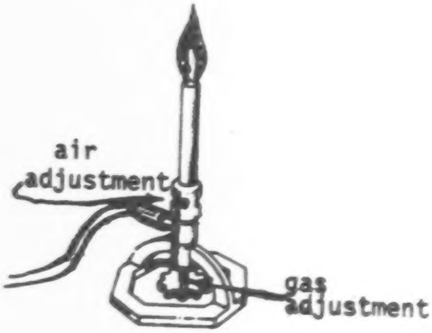
E. Test for Odor

An "educated" nose is a very useful asset in the chemistry laboratory. However, you should smell chemicals only with the greatest caution because many vapors are toxic. Never hold your nose directly over any vessel containing a chemical. Instead, very gently fan vapors toward your nose. Take care not to hit the container as you fan with your hands.



F. Using a Bunsen Burner

Always hold your face away from a burner when lighting it. If a match is used, bring it up to the oxidizing zone of the flame. Be certain the rubber hose is securely attached to the gas supply. When using a burner with both air and gas adjustment valves, the gas jet on the bench top should be fully open with all gas adjustments made at the gas valve on the bottom of the burner.



G. Making Solutions of Acids

In making aqueous solutions of liquid acids, always add the acid to the water. Never add water to acid.

H. Evaporation of Liquids

Nonflammable liquids may be evaporated from an evaporating dish with a gentle direct flame or a steam bath. Gentle boiling is more efficient and safer than rapid boiling. Flammable liquids may be evaporated using a heating mantle or steam bath. Many flammable liquids are also toxic; therefore, evaporate them only in an efficient hood.

# FIRST AID AND EMERGENCY PROCEDURES FOR FIRES

Despite the best facilities and safety procedures, an accident will occasionally occur in chemistry laboratories. A knowledge of basic first aid and emergency procedures can minimize the extent of damage to people and property. The suggestions below cover those situations most likely to occur in typical high school laboratory situations.

## First Aid

1. Burns

   Whether a burn is from a flame or a hot object, immerse the affected area in cold water (ice water if possible) as soon as possible. Pain from most burns will subside after a few minutes of this treatment. If the burn is serious, consult a physician immediately. Do not apply ointment to a serious burn.

2. Chemical Spattered in Eyes

   Get the student to an eyewash fountain with utmost haste. Rinse eyes for fifteen minutes. If a hose must be used, (be sure the flow of water is not too swift) have the student turn his head to the side so the water washes from the inside corner of the eye to the outside. After drenching, if there is any concern about the condition of the eyes, consult a physician. Remember: Bases (such as sodium hydroxide) are particularly damaging to eye tissue.

3. Chemicals Spilled or Spattered on Skin

   Wash the affected area with large quantities of water. If the spill is over a large part of the body and involves a particularly dangerous chemical (such as a strong acid or base), rush the student to a safety shower.

4. Chemicals Spilled on Clothes

   Drench the clothing with water immediately and remove as quickly as possible. A lab coat or fire blanket can always provide an emergency cover-up for the sake of modesty when a student's clothes must be removed.

5. Cuts

   For minor cuts, wash the area thoroughly, apply an antiseptic, and cover with a Band-Aid.

   For cuts which produce severe bleeding, try to control bleeding with gauze pads pressed on the area with the palm of the hand. Send for help as you administer first aid.

## Emergency Procedures for Fires

1. Clothing Fire

   If a person's clothing catches fire, guide the victim quickly (without running) to the safety shower or water hose and drench him/her. Alternatively, wrap the victim in a fire blanket to smother flames.

2. Laboratory Fire

   If a fire occurs in the laboratory, get the students out first. If the fire is minor, shut off the fuel supply if possible and use a fire extinguisher directed at the base of the fire.

   If the fire is too large to extinguish, have the fire alarm sounded for the building and call the fire department.

## CHEMICAL STORAGE AND LABELING

Each high school should have all chemicals stored in one chemical storage area supervised by a qualified person. The storage area, which should not be in a basement, should be properly marked or identified. The room should be locked when not in use and available only to authorized personnel.

The storage area should include the following safety features:

1. Two or more clearly marked exits unless the room is quite small.
2. Clean-up materials for spills.
3. Aisles free from obstruction.
4. Adequate lighting.
5. Electrical outlets and equipment must be permanently grounded.
6. Fire or smoke detectors.
7. A communication system to the main office, emergency center, or security office.
8. Fire extinguishers (multi-purpose type that covers class A, B, and C fires), one near an exit and the other at the far interior of the storage area.
9. Ideally there should be some type of respiratory protective equipment near an exit. However, inexpensive masks are not all-purpose in nature, and self-contained air systems are expensive.
10. Forced ventilation from floor to ceiling with the inlet duct located about one foot above floor level. The exhaust system should be capable of four to six changes of room air per hour. The supply air ahould be located on the opposite wall from the exhaust system.
11. Atmosphere of the room should be dry and the temperature ideally kept between 55-80°F.
12. Shelving units should be of wood construction with fixed wooden supports.
13. Avoid metal shelving units or adjustable metal shelf supports.
14. Shelving units should be securely fastened to wall or floor.
15. Shelving should have a lip.
16. No containers of chemicals should be stored above the eye level of an average-size person.

17. No containers should protrude over shelf edges.

18. Chemicals should be kept in airtight bottles.

19. No glass-stoppered bottles should be used for stockroom storage.

20. Containers should be clearly labeled as to contents, manufacturer's name and address, and lot number.

21. Labels should be legible and securely attached to containers.

22. Containers should be labeled with appropriate warning such as flammability hazard, heath hazard, reactivity, and special characteristics of which the user should be aware.

23. Containers should be labeled with date received, date opened, date for disposal of contents, and disposal method.

24. No chemicals should be exposed to direct sunlight or to a source of heat.

25. Water-sensitive chemicals should not be stored in an environment with a fire extinguishing system that utilizes water.

26. Chemicals should be stored by compatible families rather than in alphabetical order. See page 14 for a suggested arrangement of compatible chemical families as presented in School Science Laboratories -- A Guide to Some Hazardous Substances (See Safety Bibliography).

27. Incompatible chemicals should be separated from each other.

28. Flammable liquids in containers larger than one pint should be kept in approved safety cans or cabinets.*

29. Neither explosive chemicals nor carcinogens should be present for any reason in a high school storage room.

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*The National Fire Protection Association has established the following code for storing flammable and combustible liquids. (A flammable liquid has a flash point below 100°F. A combustible liquid has a flash point at or above 100°F.) No container should have a capacity greater than one gallon. Not more than ten gallons may be stored outside of a storage cabinet or storage room except in safety cans. Not more than twenty-five gallons may be stored in safety cans outside of a storage room or storage cabinet. If the quantity is twenty-five gallons or more, the liquid must be stored in an inside storage room or storage cabinet.

30. Large bottles of acids should be stored on a low shelf or in an acid cabinet. (Purchase of acids in large bottles is not recommended. Acids may be purchased in bottles coated with a plastic. Even if the bottle breaks, the liquid remains largely contained in the plastic shell.)

31. Acids should be stored away from caustic materials and active metals.

32. Acids should be stored away from chemicals that could generate toxic gases upon contact.

33. Oxidizing acids, such as nitric acid, should be stored away from organic acids, flammables, and combustible materials.

34. Solutions of inorganic hydroxides should be stored in plastic containers.

35. If a central storage facility is not available, flammable and toxic chemicals should be stored in cabinets made of wood at least one inch thick or in steel cabinets. If practical, the cabinets should be connected to a mechanical exhaust system which will help to prevent the accumulation of toxic or explosive chemical vapors.

36. Gas cylinders should be secured to prevent falling over.

37. Gas cylinders should be stored in a cool, dry place away from highly flammable substances and corrosive chemicals.

38. A perpetual chemical inventory should be kept. The inventory could be either computerized or manual.

# A SUGGESTED PATTERN FOR CHEMICAL STORAGE

Reproduced from School Science Laboratories -- A Guide to Some Hazardous Substances by the Council of State Science Supervisors. (See chemical safety bibliography at the end of this chapter.)

The alphabetical method of storing chemicals presents hazards because chemicals which react violently with each other may be stored in close proximity. The J. T. Baker Chemical Company has devised a simple color coding scheme to address this problem. The code includes both solid and striped colors which are used to designate specific hazards as follows:

Red — Flammability hazard: Store in a flammable chemical storage area.
Red Stripe — Flammability hazard: Do not store in the same area as other flammable substances.
Yellow — Reactivity hazard: Store separately from other chemicals.
Yellow Stripe — Reactivity hazard: Do not store with other yellow coded chemicals; store separately.
White — Contact hazard: Store separately in a corrosion-proof location.
White Stripe — Contact hazard: Not compatible with chemicals in solid white category.
Blue — Health hazard: Store in a secure poison area.
Orange — Not suitably characterized by any of the foregoing categories.

Once the chemicals are sorted according to ther color codes, sorting into organic and inorganic classes within a color should occur. The Flinn Chemical Catalog Reference Manual suggests organic and inorganic groupings which are further sorted into compatible families. The compatible families suggested* are:

**INORGANIC**

1. Metals, hydrides
2. Halides, sulfates, sulfites, thiosulfates, Phosphates, halogens
3. Amides, nitrates** (except ammonium nitrate), nitrites**, azides**, nitric acid
4. Hydroxides, oxides, silicates, carbonates, carbon
5. Sulfides, selenides, phosphides, carbides, nitrides
6. Chlorates, perchlorates**, perchloric acid**, chlorites, hypochlorites, peroxides**, hydrogen peroxide
7. Arsenates, cyanides, cyanates
8. Borates, chromates, manganates, permanganates
9. Acids (except nitric)
10. Sulfur, phosphorus**, arsenic, phosphorus pentoxide**

**ORGANIC**

1. Acids, Anhydrides, peracids
2. Alcohols, glycols, amines, amides, imines, imides
3. Hydrocarbons, esters, aldehydes
4. Ethers**, ketones, ketenes, halogenated hydrocarbons, ethylene oxide
5. Epoxy compounds, isocyanates
6. Peroxides, hydroperoxides, azides**
7. Sulfides, polysulfides, sulfoxides, nitriles
8. Phenols, cresols

Using a combination of the J. T. Baker and Flinn Scientific storage schemes should eliminate chemical incompatibilities in the chemical storage room.

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**These chemicals deserve special attention due to their potential instability.

## SAFETY CHECK LIST

## CHEMISTRY LABORATORY

Use this list as a guide to evaluate the safety of your laboratory. The list does not cover every conceivable safety item but certainly some of the most important items are covered.

YES NO

( ) ( ) 1. Does the room have two clearly marked exits?
( ) ( ) 2. Is the laboratory clean with the aisles free of obstruction?
( ) ( ) 3. Is a fire or smoke detector installed?
( ) ( ) 4. Does the detector function?
( ) ( ) 5. Is there a fire extinguisher near door?
( ) ( ) 6. Is there a fire extinguisher near the center or back of the room?
( ) ( ) 7. Are the fire extinguishers regularly inspected?
( ) ( ) 8. Is there a fume hood for use during dangerous experiments?
( ) ( ) 9. Are electrical outlets and equipment grounded and in good repair?
( ) ( ) 10. Is there a safety shower?
( ) ( ) 11. Is there an eyewash fountain or hose?
( ) ( ) 12. Are first aid supplies readily available?
( ) ( ) 13. Does an emergency warning / evacuation procedure exist?
( ) ( ) 14. Is there personal body protection (aprons, gloves) available?
( ) ( ) 15. Is there proper eye protection available and worn by everyone?
( ) ( ) 16. Are contact lenses prohibited from being worn in the lab?
( ) ( ) 17. Is all glass apparatus made of borosilicate glass?

## SAFETY CHECK LIST

## CHEMISTRY STOCKROOM

Use this list as guide to evaluate the safety of your stockroom. The list does not cover every conceivable safety item but certainly some of the most important items are covered.

YES NO

( ) ( ) 1. Has a stockroom inventory been completed?

( ) ( ) 2. Have all explosive chemicals been disposed of properly?

( ) ( ) 3. Have all toxic and known carcinogens been removed?

( ) ( ) 4. Do shelves have a lip to prevent bottles from sliding off?

( ) ( ) 5. Are shelves secure on walls?

( ) ( ) 6. Are there two exits from the chemical storeroom?

( ) ( ) 7. Is there an exhaust ventilation system to remove vapors?

( ) ( ) 8. Is there a fire extinguisher near door of storeroom?

( ) ( ) 9. Is there a chemical spill kit near door of storeroom?

( ) ( ) 10. Are chemicals arranged in a manner that separates acids, bases, oxidizers, etc., such as via the "FLINN" or the "BAKER" system?

( ) ( ) 11. Are flammables stored in a cabinet made of metal or 1-inch wood?

( ) ( ) 12. Are large bottles stored on low shelves and nothing stored above eye level?

( ) ( ) 13. Are there labels on all purchased chemicals indicating date received and special precautions?

( ) ( ) 14. Are locally-prepared solutions labeled with name, date prepared, concentration and initials of person preparing?

( ) ( ) 15. Are all labels covered with clear tape or plastic to prevent fading or removal by fumes and humidity?

# LISTS OF DANGEROUS CHEMICALS

(These lists are not comprehensive. They only include some of the more common dangerous chemicals.)

The substances listed below in categories I, II, and III should NOT be used or stored in high schools. If present they should be removed as soon as possible using the safest removal techniques. After removal they should NOT be reordered.

Most chemicals in category IV are rarely used in high school experiments and should be permanently removed. However, barium hydroxide, mercury, and mercuric salts do find frequent use. If you decide to retain these members of category IV, be aware of their extreme toxicity and take special precautions when they are used.

## I. EXPLOSIVE CHEMICALS OR THOSE THAT FORM EXPLOSIVE PRODUCTS ON STANDING OR IN USE

benzoyl peroxide
carbon disulfide
diisopropyl ether
ethyl ether
picric acid
perchloric acid
potassium metal

Since some of these chemicals could conceivably detonate with little or no movement when old or dry, do NOT remove lids from containers. If possible, avoid moving the container at all. Notify your principal immediately if these chemicals are present. Get expert advice and assistance before removing these containers.

## II. KNOWN CARCINOGENS

arsenic powder
arsenic pentoxide
arsenic trichloride
arsenic trioxide
asbestos
benzene
benzidine
chromium powder
chromium(VI) oxide
lead arsenate
sodium arsenate
sodium arsenite

## III. PROBABLE CARCINOGENS

- cadmium powder
- cadmium chloride
- cadium sulfate
- carbon tetrachloride
- chloroform
- ethylene oxide
- nickel powder
- *o*-toluidine

## IV. HIGHLY TOXIC CHEMICALS

- adrenaline
- barium hydroxide
- chlorine
- mercury[*]
- mercuric chloride
- mercuric iodide
- mercuric nitrate
- mercuric oxide
- mercuric sulfate
- nicotine
- osmium tetroxide
- phosphorus (white)
- phosphorus pentoxide
- potassium cyanide
- potassium periodate
- silver cyanide
- sodium cyanide

(Very small amounts of the chemicals in Group IV may cause immediate toxic reactions.)

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[*] When thermometers are ordered, consider purchasing the non-mercury type if the graduations, markings, etc. are appropriate for your needs.

# UNSAFE LABORATORY EXPERIMENTS

Presented below is a list of some classical laboratory experiments or reactions that are now known to present significant health or safety concerns. Because these experiments are hazardous, they should not be included in a high school laboratory curriculum.

1. Generation of oxygen from potassium chlorate (explosion hazard).
2. Heating or evaporating organic solvents using open flame (fire hazard).
3. "The thermite reaction" (too violent).
4. Decomposition of mercury(II) oxide (generates toxic mercury vapor).
5. Ammonium dichromate-organic solvent "volcanoes" (toxic $Cr^{+6}$ ions and violent reaction).
6. Group IA metal reactions with water (violent reaction).
7. Sublimation or heating of iodine (toxic iodine vapor).
8. Use of chromic acid cleaning solutions (toxic; extremely corrosive; $Cr^{+6}$ is carcinogenic).
9. Qualitative analysis experiments using hydrogen sulfide gas (extremely toxic; may temporarily paralyze olfactory nerve.)
10. Use of carbon disulfide (extraordinarily flammable).

## DISPOSAL OF HAZARDOUS SUBSTANCES

Quite often the high school science teacher inherits a stockroom that contains hazardous substances. This situation has many potential origins:

(1) Chemicals have been ordered over the years for experiments performed in the past but now considered too dangerous for the high school laboratory.

(2) Gifts of chemicals from industry or other sources.

(3) Acquisition of free chemicals from U. S. government surplus warehouses.

(4) Decomposition of chemicals over the years.

(5) Reclassification (by authorities) as "hazardous" of chemicals once believed to be safe.

It can be said with reasonable certainly that most high school chemistry stockrooms contain some chemicals considered too hazardous to be stored there. This situation leads to a disposal problem for the high school chemistry teacher.

On pages 17 and 18 of this chapter you will find lists of extremely hazardous substances. If funds are available, you should dispose of dangerous substances by employing a professional chemical waste disposal company. However, if you DO have hazardous substances and you DON'T have funds to pay for their removal, you are faced with three basic options:

(1) Do nothing.

The "it won't happen to me" syndrome is common but potentially fatal Some chemicals such as aged picric acid and peroxides (including those formed in ethers) are literally time bombs waiting to explode. (If you have either of these substances in your laboratory, you should not attempt to dispose of them without professional help.) Keep in mind the fact that if a hazard is known to exist by persons in positions of responsibility and nothing is done by them to remove that hazard, the liability of these individuals is very great.

(2) Dispose of substances in a haphazard way.

As a teacher of science, you cannot afford to callously disregard the safety of others and the environment despite a convenient opportunity to do so or any coercion that might be applied for use of this method.

(3) Solve the problem by making it a learning experience.

Turn a troublesome situation into a learning situation by following some logical, systematic steps:

(a) Define the problem.

Inventory your stockroom thoroughly and determine which substances are hazardous.

(b) Investigate disposal methods.

Before any steps are taken to actually dispose of a substance, a proper, safe method must be found.

Mississippi high schools should begin the quest for information by telephoning The Mississippi Department of Natural Resources (Phone 601-961-5099) and asking for the extension of the Division of Solid Waste Management.

Sources of information available to residents of all states include:

(i) University Chemistry Departments

The use of hazardous chemicals is much more common in universities than in high schools because of the research function of a university. There should be persons in a large chemistry department who are aquainted with the specific properties of the substance in question and may have already developed a specific and tested plan of disposal. Keep in mind, however, that university faculty are extremely busy and typically cannot take on your problem as their own because of lack of time. Often they can refer you to good sources of help, however.

(ii) Governmental Agencies (OSHA, EPA) and Local Agencies (Fire Department, Civil Defense, State Department of Health)

Methods of disposal are often regulated by law. One or more of these agencies should be contacted in order to ensure that these laws are followed. They will possibly know of skilled individuals that may be able to assist in the actual disposal.

(iii) Resource Materials

There are several texts that have been written about proper disposal of hazardous substances. Some of the best are:

Prudent Practices for Disposal of Chemicals from Laboratories by the National Research Council Committee on Hazardous Substances in the Laboratory; 1983; National Academy Press.

A modern text, with excellent ideas and suggestions.

The Care, Handling and Disposal of Dangerous Chemicals; P. J. Gaston; The Institute of Science Technology; Northern Publishers Limited; Aberdeen, Scotland; 1970.

An older text but with much specific detail concerning dangerous substances and their disposal. Its suggestions should be reviewed to determine compatibility with EPA regulations.

Chemical Catalog/Reference Manual, Flinn Scientific, Inc.; Box 231; Batavia, IL 60510.

An excellent safety source that should be in all labs. It is weaker than the two previous texts in the area of disposal in that it deals with groups of chemicals generally rather than specific chemicals.

(c) Decide if there is a safe method of disposal.

After all sources of information have been reviewed you should be able to decide if there is a safe method of disposal. If not, you may have no other alternative than to have it done professionally. If this is the case, school officials must locate funds for this purpose. If there is a safe method of disposal, develop the procedures that you will use to dispose of the substance. Document your methods, materials to be used, procedures to be followed and sources. Then have your procedure reviewed by someone with expertise in the area. It would be wise to have only competent, trained persons assist you in the actual disposal.

(d) Make your methods available to others for information and review.

Your work in developing a safe and tested method will be appreciated by your fellow teachers.

By following a logical, methodical approach to this problem, you may be able to defuse a hazardous situation in your school, broaden your knowledge in chemical safety and disposal, and illustrate to your students that problems may be solved if attacked logically.

## THE USE OF SAFETY RULES AND STUDENT CONTRACTS

The use of a list of safety rules and a "safety contract" can be very effective in increasing students' and parents' awareness of safety in the science laboratory. A contract between the student and teacher which has been read and signed, not only by the student, but by the student's parents can be a valuable tool in promoting desirable student behavior in the laboratory. Often students and parents are unaware of the hazards that accompany laboratory work. Therefore, it is important that the teacher point out known hazards and how to look for others. This approach should insure cooperation; from the students as well as the parents.

The contract document should include as a supplement a list of the most important safety rules to be followed in the laboratory. Parents thus will learn of the rules which their children are asked to follow. By implication the dangers of disregarding the rules are obvious.

A good list of safety rules will be comprehensive, but cannot be too exhaustive. Students are easily "turned off" by a long list of "do nots". Also, a positive approach will usually be more effective in obtaining a desirable response.

The safety rules should be posted at appropriate locations in the classroom, laboratory, and stockroom to insure the students' attention.

Examples of a list of safety rules and a safety contract are presented on the following pages. Only some of the most important rules are listed in this example. It is suggested that each teacher add those rules which he/she considers important.

## CHEMISTRY LABORATORY SAFETY RULES FOR STUDENTS

1. Do only the experiment assigned or approved by the teacher. Unauthorized experiments are prohibited.
2. Wear proper eye protection and laboratory aprons or other protective clothing during all laboratory activities and in locations where chemicals are sorted or handled.
3. Do not smoke, eat, drink, or chew gum or tobacco in the laboratory. Dangerous chemicals may get in the mouth or lungs!
4. Never engage in horseplay or practical jokes.
5. Dress appropriately on days when laboratory work will be done. Shorts, sandals, or loose and baggy clothes should not be worn.
6. Confine or securely tie hair that reaches the shoulder. Remember that hair is flammable.
7. Do not wear contact lenses in the laboratory.
8. Never do any procedure that you are unsure of.
9. Immediately report all accidents, no matter how minor, to the instructor.

# STUDENT SAFETY CONTRACT

I, ______________________________, a student in chemistry at ____________________High School, do hereby agree to follow all safety rules and regulations as set forth by the instructor. I realize that compliance with these rules is necessary to assure the safe operation of the school laboratory and provide a safe environment not only for myself, but for my fellow students and teachers as well. I will, therefore, cooperate fully with the teacher and students to assure all of us the safest laboratory possible. I will act responsibly to look for possible safety hazards, and will immediately point out these hazards to the instructor. I realize that, as a student, much of the responsibility for safety is in my hands. I have read the attached rules and will comply with them to the best of my ability. I understand that violation of these rules may result in the loss of laboratory privileges and possible disciplinary measures.

Signed:________________________________Student

Date:__________________________

I, ____________________________________, parent (or guardian) of ___________________________________ have read this safety contract and the attached rules list. I give my permission for my child (or ward) to enroll in this chemistry course.

Signed: _______________________________Parent (or Guardian)

Date:_________________________

# CHEMICAL SAFETY BOOKS OF INTEREST TO CHEMISTRY TEACHERS

1. Safety in Academic Laboratories, 4th Edition, American Chemical Society, 1985. 1st Copy FREE, addnl. copies $1.00. ISBN 0-8412-0938-3

   Order From: American Chemical Society
   Distribution Office, Dept. 22H
   1155 Sixteenth St. N.W.
   Washington, D.C. 20036

2. Safety in the School Science Laboratory, Instructor's Resource Guide, National Institute for Occupational Safety and Health, 1979. Copies FREE

   Order From: National Institute for Occupational Safety and Health
   Division of Training and Manpower Development
   Cincinnati. Ohio 45226

3. Manual of Safety and Health Hazards in the School Science Laboratory, National Institute for Occupational Safety and Health, 1980. $5.75

   Order From: Council of State Science Supervisors
   Route 2, Box 637
   Lancaster, VA 22503

4. School Science Laboratories, A Guide to Some Hazardous Substances (A Supplement to Manual of Safety and Health Hazards in the School Science Laboratory), Council of State Science Supervisors, 1984. Sent Free with book above.

   Order From: Council of State Science Supervisors
   Route 2, Box 637
   Lancaster, VA 22503

5. Chemical Catalog/Reference Manual, Flinn Scientific, Inc. FREE

   Order From: Chemical Catalog/Reference Manual
   Flinn Scientific, Inc.
   P.O. Box 231, 917 W. Wilson Street.
   Batavia, IL 60510

6. Safety in the Secondary Science Classroom, National Science Teachers Association, 1978. $4.00

   Order From: National Science Teachers Association
   1742 Connecticut Avenue N.W.
   Washington, D.C. 20009

Chemical Safety Books
of Interest to Chemistry Teachers, page 2

7. Safety in Working with Chemicals, by Michael E. Green and Amos Turk, Macmillan Publishing Co., 1978. Paperback. About $12.00. ISBN 0-02-346420-8

Order From: Macmillan Publishing Co.
866 Third Avenue
New York, NY 10022

8. Prudent Practices for Handling Hazardous Chemicals in Laboratories, National Academy Press, 1981. Paper Text Edition. $9.95. ISBN 0-309-03128-1

Order From: National Academy Press
2101 Constitution Avenue, N.W.
Washington, D.C. 20418

9. Prudent Practices for Disposal of Chemicals From Laboratories, National Academy Press, 1983. $16.50. ISBN 0-309-03390-X

Order From: National Academy Press
2101 Constitution Avenue, N.W.
Washington, D.C. 20418

10. School Science Safety: Secondary, Jack A. Gerlovich, Editor, Flinn Scientific, Inc., 1984. $15.95.

Order From: Flinn Scientific, Inc.
P.O. Box 231, 971 W. Wilson St.
Batavia, IL 60510

11. Dangerous Properties of Industrial Materials, 5th Edition, by N. Irving Sax, Van Nostrand-Reinhold, 1979. $57.50. ISBN 0-442-27373-8

Order From: Van Nostrand Reinhold Co.
Division of Litton Educational Publishing
135 W. 50th Street
New York, NY 10020

12. Handbook of Laboratory Safety, 2nd Edition, Norman V. Steere, Editor, CRC Press, 1971. $59.95. ISBN 0-8493-03524

Order From: CRC Press
2000 Corporate Blvd.
Boca Raton, FL 33431

Chemical Safety Books
of Interest to Chemistry Teachers, page 3

13. The Merck Index, 10th Edition, Merck and Co., Inc., 1983. $28.50.
ISBN 0-911910-26-3

Order From: Merck and Co., Inc.
Box 2000
Rahway, NJ 07065

14. Safety Equipment Catalog, Lab Safety Supply Co. FREE

Order From: Lab Safety Supply
Division of Science Related Materials
P.O. Box 1368
Janesville, Wisconsin 53547-1368

15. Safety Equipment Catalog, Direct Safety Co. FREE

Order From: Direct Safety Company
7815 South 46th Street
Phoenix, AZ 85040

## A GENERAL METHOD FOR DISPOSING OF SOLUTIONS OF CERTAIN TOXIC INORGANIC SALTS INCLUDING THOSE OF BARIUM, LEAD, AND SILVER

Note: This procedure is based upon forming an insoluble carbonate salt of the toxic cation. You cannot use this procedure with just any toxic inorganic salt. Do not attempt the procedure with compounds involving cations that form soluble carbonates or hydroxides. If you are not 100% certain of the chemistry involved, don't use the procedure.

Precaution: CARRY OUT THIS PROCEDURE EITHER IN A GOOD HOOD OR OUTDOORS. WEAR SAFETY GOGGLES, GLOVES, AND A LABORATORY COAT OR APRON.)

Fill a plastic bucket of appropriate size about half full of tap water. Slowly add the inorganic salt to be discarded while stirring with a wooden stick or paddle. Next, slowly add about four times as much sodium carbonate* (soda ash) as the weight of inorganic salt. Allow the mixture to stand undisturbed in a safe location for 24 hours. Solid material will settle to the bottom of the bucket with a liquid layer covering it. Very carefully pour off the liquid into another container. Neutralize this liquid (in a hood or outdoors) with 6 molar hydrochloric acid solution to pH 7 or slightly above using pH indicator paper. Wash the neutralized solution down a sink drain with a large quantity of water. The water left in the sludge remaining in the plastic bucket can be allowed to evaporate. The dried residue is then removed from the bucket, placed in a heavy paper bag, and discarded at a dump with the approval of the operator of the dump. Carefully rinse the bucket clean.

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*Use the cheapest grade available.

## AN UPDATE ON THE DISPOSAL OF HAZARDOUS CHEMICALS -- SEPTEMBER 1985

In some states and localities, local laws and regulations concerning disposal of small amounts of chemical waste are not specific enough and possibly inadequate. In Mississippi, for example, it appears at this time that if a chemical is not listed on a special list of extremely toxic substances, a small amount of the substance may be discarded at a local dump with permission of the owner or operator. Don't let this potential "easy way out" tempt you to forget common sense. No chemicals should be discarded in a local dump if there is the slightest chance that they could injure the operator or other personnel who might visit the dump site, or if water run-off following a rain would carry them into a nearby stream.

The problem of chemical waste disposal is exceedingly complex. Laws are changing rapidly and it is difficult to reduce the problem to a set of simple procedures. The removal of dangerous chemicals from the high schools of Mississippi is a high priority objective of the Education Committee of the Mississippi Section of the American Chemical Society. As helpful information about the subject becomes available, the Committee plans to keep the high school chemistry teachers informed.

If, in your judgement, it is absolutely impractical to remove all hazardous chemicals from your stockroom immediately, isolate the non-explosive, hazardous chemicals, separate them by category, and store them in a safe and secure location until you get more specific instructions. Explosive substances must be removed immediately, but only with the aid of experts.

# CHAPTER 6

# CHEMISTRY TEACHERS AND THE LAW

# CHAPTER 6 -- CHEMISTRY TEACHERS AND THE LAW

## INTRODUCTION

The legal responsibility for any accident in the laboratory has the potential for being placed directly on the teacher's shoulders, regardless of his/her awareness of this responsibility. All teachers have the responsibilities of instruction and supervision, but the chemistry teacher has the added responsibility of maintenance regarding chemicals and facilities.

The teacher must be aware of possible hazards in the experiment, and is legally bound to inform the students of these possible hazards and the precautions to be taken to avoid injury. The instructor's ignorance of hazardous conditions will not absolve him/her from liability due to negligence, which is the omission of reasonable precautions. The laboratory instructor must be very careful to protect him/herself against charges of negligence.

Read again the guidelines for the laboratory instructor found on page 2 of the chapter on Chemical Safety. Following these guidelines is the beginning of a program of self defense against some future charge of negligence.

By using the safety contract form to be signed by students and parents, the teacher makes both students and parents aware of possible hazards and the absolute necessity of enforcing safety rules in the laboratory.

Too often the chemistry teacher is put in the untenable position of having to teach in an unsafe environment which he/she has no resources to improve. The logical recourse in such a situation is to inform the administration of the unsafe situation and request that the situation be remedied. It is believed that once the request has been made via a formal, written document, liability that once was exclusively the teacher's is shared or possibly totally assumed by the individual administrators.

In order to assist teachers in getting safety problems corrected, three "example letters" are presented on the concluding pages of this chapter. These letters, which differ in subject matter and format, may serve as models of letters which teachers can write to address their own special situations. It is certainly advisable to first talk with administrators about the problems and let them know that you plan to write them in detail about the matter. Certainly maintaining the good will of the administration is of utmost importance in getting the resources needed to improve safety in the laboratory.

An excellent, very readable discussion of the legal considerations involved in laboratory instruction can be found in School Science Safety--Secondary by Gerlovich, Gerard, Shriver, Downs and Flinn, Flinn Scientific, Inc., Batavia, Illinois, 1984. The first chapter of this book should be required reading for any high school chemistry teacher.

## EXAMPLE LETTER 1 -- REMOVAL OF HAZARDOUS CHEMICALS

October 1, 1985

Mr. A. B. Jones
Principal, Smith High School

Dear Mr. Jones:

As educators, we share a common concern for school safety. Our science storeroom currently contains several hazardous chemicals (listed below) which should be removed immediately. The enclosed reference materials verify that these chemicals constitute a sufficiently great health and/or safety hazard to warrant your immediate attention. The science department does not have the resources to dispose of this material in accordance with federal regulations; therefore, it is necessary to seek action on your part. Since there are very specific regulations governing chemical waste disposal (with severe penalties for violation), it is important that appropriate authorities be consulted prior to any action being taken by any personnel in our school system. Please give this matter your prompt attention because these chemicals constitute a genuine, immediate safety and health concern.

[List chemicals here.]

Respectfully submitted,

[Sign your name.]

D. C. Brown
Chemistry Teacher

ENCLOSURE: Reference Materials

## EXAMPLE LETTER 2 -- NEED FOR SAFETY SUPPLIES AND EQUIPMENT

October 1, 1985

### MEMO

TO: Mr. A. B. Jones
Principal, Smith High School

FROM: D. C. Brown, Chemistry Teacher

SUBJECT: Laboratory Safety Equipment/Supplies

Of growing concern to me is the safety of the students in our science laboratories. I am sure that you and the general public share this concern. Since court cases have shown that you and I, as educators, are legally liable for students' safety and health, I am respectfully requesting your prompt action in correcting the safety deficiencies mentioned herein.

I would like to call to your attention the fact that the safety equipment and supplies listed below are recommended by safety authorities [such as those at the National Institute of Safety and Health (NIOSH)] but are not found in our school's laboratory.

| | Item | Approximate Cost |
|---|---|---|
| 1. | | |
| 2. | | |
| 3. | [List items and cost here.] | |
| 4. | | |
| 5. | | |

I personally feel that these items are essential to the maintenance of a safe laboratory environment, and I urge you to see that they are provided as soon as possible.

Respectfully yours,

[Sign your name.]

## EXAMPLE LETTER 3 -- SAFETY DEFICIENCIES IN THE SCIENCE LABORATORY

October 1, 1985

Mr. A. B. Jones
Principal, Smith High School

Dear Mr. Jones:

I am writing to you in keeping with our commitment to promote safety in the school and in the science laboratory. It has come to my attention through discussion with safety experts in the area that ________(room - laboratory - etc.)________ used for the teaching of ________(class)________ is deficient in the following respects. Although other deficiencies may well exist at this time, I feel that these are the most critical and need to be remedied immediately.

[List the deficiencies in order of priority from most critical to least critical. Include here reasons for the needs. Be explicit!]

To fulfill our legal, moral, ethical and professional responsibilities as science educators, to ourselves and our students, we must begin the process of correcting the aforementioned inadequacies. I respectfully request your assistance.

Sincerely,

[Sign your name.]

D. C. Brown
Chemistry Teacher

[This letter should be used for the most critical needs. Financially unreasonable requests such as $10,000 worth of safety cabinets will only be counter-productive to getting immediate action on critically important issues.]

## TWELVE STEPS TOWARD FREEDOM FROM WORRY ABOUT NEGLIGENCE LAWSUITS

1. Using a checklist, conduct a safety survey and hazard analysis of your laboratory and stockroom.

2. Do whatever you can to make the laboratory room and stockroom safer.

3. Inform your principal in writing (keeping a copy) of hazards, deficiencies, need for removal of hazardous chemicals, etc. which will require action by him/her or a higher-level administrator.

4. Reevaluate all laboratory experiments and modify or eliminate those that appear to be unsafe.

5. Develop safety rules jointly with your principal for the dress and behavior of students in the chemistry laboratory. Jointly agree on the penalties for violations of these rules.

6. Issue the printed safety rules to students and admit them to your chemistry class only if they sign a statement that they will abide by these rules (constituting an informal safety contract). File these signed statements.

7. Send a letter to the parents or guardian of each student stating that (1) there is some element of hazard involved in being a student in a chemistry laboratory; (2) you will conduct the laboratory as safely as you can; (3) you will rigidly enforce safety rules which will be made known to all students; (4) penalties for disobedience of rules will be severe. Require the parents (guardian) to sign a statement saying that they give permission for their child to take the chemistry course. Require this signed statment of parents (in addition to the student's signed statement) for admission of the student to your chemistry class. File these signed statements.

8. Give the best safety training you can to your students with general safety sessions at the beginning of the course and specific safety directions with each experiment. Emphasize safety by your personal example and by your continuing references to the safe way to perform various laboratory operations.

9. When students encounter both a hazardous chemical and a new technique in the same experiment, if possible have them practice and develop their technique with a harmless chemical prior to using the hazardous chemical. (For example, practice measuring volumes in a graduated cylinder using water rather than sulfuric acid.)

10. Maintain good stockroom records.

11. Develop an accident report form to be completed and signed by a student who is injured in the laboratory. File these.

12. Purchase professional liability insurance. (Such insurance is often available via a "rider" to a homeowner's insurance policy.)

# CHAPTER 7

# DEMONSTRATIONS

# CHAPTER 7 -- DEMONSTRATIONS

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## DEMONSTRATION 1

## COMMON HOUSEHOLD CHEMICALS ON DISPLAY

### GENERAL TOPIC: I. INTRODUCTION

Materials:

baking soda
detergents
salt
toilet bowl cleaner
vinegar
water

Procedure: Place commercial containers of common household chemicals on a table. Explain that all these materials represent practical uses of chemicals and chemical principles. Make students aware of chemical information found on the labels of the containers.

Hazards: None, if normal care is used in handling of materials, and if the container of toilet bowl cleaner (which is a hazardous material) is not opened.

Questions:

1. Name some chemicals used in food preparation.

   Answer: water, salt, vinegar, baking soda, many others

2. What types of chemicals are used in the bathroom?

   Answer: soaps, detergents, toilet bowl cleaners, tile cleaner, others

## DEMONSTRATION 2

## "POWERS OF TEN" (A 16 mm Film)

### GENERAL TOPIC: III. SCIENTIFIC NOTATION AND SIGNIFICANT FIGURES

There is an excellent film that can be rented which covers scientific notation.

TITLE: "Powers of Ten"

BY: Herman Miller

LENGTH OF FILM: Less than 30 minutes

ADDRESS: Resources Center Film Library
8500 Byron Road
Zeeland, MI 49464

COST: $15 for 5 days

## DEMONSTRATION 3

## NAMES, SYMBOLS AND FORMULAS OF COMMON CHEMICALS

## GENERAL TOPIC: IV. NAMING, SYMBOLS AND FORMULAS

**Materials:**

| Common Name or Form | Chemical Name | Symbol or Formula |
|---|---|---|
| aluminum foil | aluminum | Al |
| calcium chloride | calcium chloride | $CaCl_2$ |
| chalk | calcium carbonate | $CaCO_3$ |
| pencil lead | carbon | C |
| potassium chloride | potassium chloride | KCl |
| rust | ferric oxide or iron(III) oxide | $Fe_2O_3$ |
| salt | sodium chloride | NaCl |
| sand | silicon dioxide | $SiO_2$ |
| sugar | sucrose | $C_{12}H_{22}O_{11}$ |
| vinegar | acetic acid | $CH_3COOH$ |
| water | water | $H_2O$ |

**Procedure:** Place a collection of common chemicals on a table. Discuss the elemental composition and formula of each. Consider different states of matter.

**Hazards:** None

**Questions:**

1. What is the formula of sand?

   Answer: $SiO_2$

2. What is the formula of water?

   Answer: $H_2O$

3. What element is in pencil lead?

   Answer: Carbon

## DEMONSTRATION 4

## COMBUSTION REACTION

### GENERAL TOPIC: V. EQUATIONS

Materials: Bunsen burner burning natural gas (methane), $CH_4$ or butane lighter burning butane, $C_4H_{10}$

piece of metal or glass

tongs

Procedure:

1. Light a butane lighter or Bunsen burner.
2. Hold a cold piece of metal or glass above the flame and allow water to condense on the object.
3. Write the chemical equation for the combustion reaction on the chalkboard.

Butane (1st attempt):

$$C_4H_{10}(g) + \underline{\ ?\ }\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$$

(The odd number coefficient for $H_2O$ would not result in a whole number of $O_2$ molecules. So double the numbers.)

Butane (2nd attempt):

$$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$$

Methane: The methane equation is much more simple.

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

To Be Emphasized:

1. Combustion can be represented by a chemical equation.
2. Water, a product of combustion, can actually be observed as condensation on a cold surface.
3. The equation(s) represent something familiar to the students.
4. The butane combustion involves a chemical equation that requires more than one attempt to balance.

Demonstration 4
Combustion Reaction, page 2

Questions:

1. Why can the equation not be balanced with 5 $H_2O$ in the butane equation?

   Answer: Elemental oxygen exists as $O_2$ and would not be balanced with a whole number. We cannot use 6 1/2 as the coefficient, and we cannot change subscripts in chemical formulas!

2. How many molecules are there in the products of the balanced butane equation?

   Answer: 18 molecules

3. How many reactant molecules are there in the balanced butane equation?

   Answer: 15 molecules

4. Does the number of reactant molecules equal the number of product molecules in a balanced equation?

   Answer: Not necessarily

5. Does the number of atoms of reactants equal the number of atoms of products in a balanced equation?

   Answer: Yes

## DEMONSTRATION 5

## MOLAR QUANTITIES

### GENERAL TOPIC: VI. STOICHIOMETRY AND THE MOLE CONCEPT

**Materials:**

- copper
- glycerol (glycerine)
- sodium chloride
- sugar
- sulfur
- water

**Note:** All solids should be finely devided and flow freely.

**Procedure:**

1. Weigh out 18 grams of water (the mass of 1 mole of water) into a 250 mL bottle and label as:

   1 mole $H_2O$
   18 g $H_2O$
   or
   6 X $10^{23}$
   molecules $H_2O$

2. Repeat #1 using copper (63.5 g), glycerine (92.1 g), sodium chloride (58.5 g), sugar (342 g), and sulfur (32 g); weighing out 1 molar mass of each.
3. Show these to your students. Let them hold a mole.

**To Be Emphasized:**

1. One mole of an element or compound has a convenient mass and volume to work with.
2. The mass of one mole of a pure substance is proportional to the atomic or molecular mass of the substance.
3. One mole of any substance contains the same number of molecules ($6.0 \times 10^{23}$) as a mole of anything else.

**Question:** What is the common quantity in one mole of any substance?

Answer: $6.0 \times 10^{23}$ molecules

## DEMONSTRATION 6

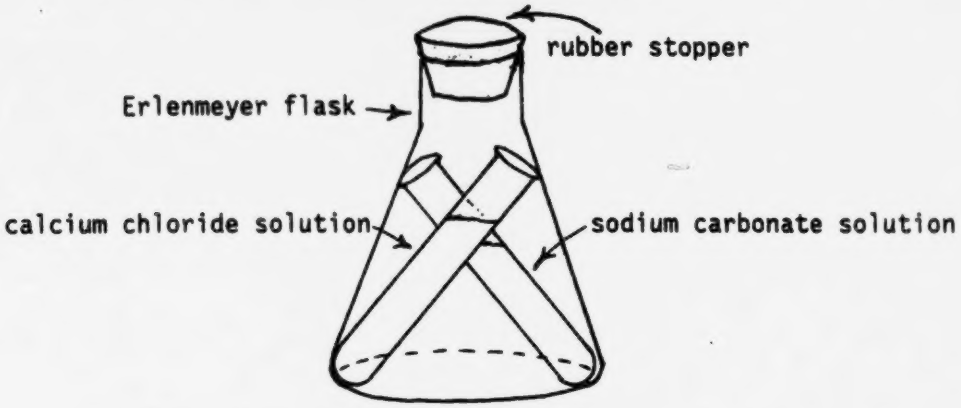
## CONSERVATION OF MASS

### GENERAL TOPIC: VI. STOICHIOMETRY AND THE MOLE CONCEPT

**Materials:**

25.0 mL of 1.0 M aqueous solution of calcium chloride, $CaCl_2$
25.0 mL of 1.0 M aqueous solution of sodium carbonate, $Na_2CO_3$
balance, 0.01 g sensitivity
250 mL Erlenmeyer flask
rubber stopper to fit the flask
two 13 x 100 mm test tubes

**Procedure:** Fill the test tubes about 2/3 full, one with $CaCl_2$ solution and the other with $Na_2CO_3$ solution. Carefully place these in the flask as shown. Stopper the flask.



Weigh the flask and contents and have the mass recorded on the chalkboard.

Hold the flask up and invert it to react the $CaCl_2(aq)$ and $Na_2CO_3(aq)$. A white precipitate will immediately form. Then weigh the flask again and note that although a chemical reaction occurred, no change in mass occurred.

**To Be Emphasized:** This clearly shows that the mass doesn't change in a chemical reaction.

**Questions:** 1. What is the reaction that occurred?

Answer: $CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$

Demonstration 6
Conservation of Mass, page 2

2. How many moles of reactants are in this equation?

   Answer: 2 moles; 1 of $CaCl_2$ and 1 of $Na_2CO_3$

3. How many moles of products?

   Answer: 3 moles. Therefore, the moles of reactants do not equal the moles of products in this reaction.

4. Is the total mass of reactants equal to the total mass of products?

   Answer: yes

## DEMONSTRATION 7

## ELEMENTS AND COMPOUNDS

## GENERAL TOPIC: VII. DESCRIPTIVE CHEMISTRY

Materials:

acetone
aluminum (bolt from hardware store)
barium chloride, $BaCl_2 \cdot 2H_2O$
benzoic acid
bismuth (lump)
calcium chloride, $CaCl_2 \cdot 2H_2O$
camphor
chlorine gas (produced on site; see Procedure. Also see comments under Hazards)
"Clorox"
cobalt(II) nitrate hexahydrate, $Co(NO_3)_2 \cdot 6H_2O$
copper(II) acetate monohydrate, $Cu(C_2H_3O_2)_2 \cdot H_2O$
copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$
ethyl alcohol
concentrated hydrochloric acid, HCl
magnesium ribbon
magnesium chloride hexahydrate, $MgCl_2 \cdot 6H_2O$
napthalene
phosphorus, red (See comments under Hazards.)
potassium permanganate, $KMnO_4$
1 M aqueous solution of sodium hydroxide, NaOH
sugar
sulfur
water
250 mL Erlenmeyer flask
graduated cylinder
containers

Procedure:

(A) Show samples of magnesium ribbon (a solid), aluminum (solid), sulfur (solid), and chlorine (gas). SEE COMMENTS UNDER "HAZARD".

(Chlorine gas can be made by adding 1.0 mL of concentrated hydrochloric acid to 25 mL of "Clorox" in a 250 mL Erlenmeyer flask. This reaction can be stopped by the addition of 15 mL of 1 M NaOH solution.)

Demonstration 7
Elements and Compounds, page 2

Explain to the students the correlation between the characteristics of the elements and the location of the symbols of the elements on the periodic chart. You should point out the change in appearance, physical state, and color of the elements as you refer to their symbols from left to right in Period 3 of the periodic chart. Magnesium is a solid metal; sulfur is a solid but is a yellow powder; and chlorine is a greenish-yellow gas. These trends are shown in going from metallic to nonmetallic elements.

(B) A similar trend may be shown with Group V elements. Everyone recognizes nitrogen as a gas. Show a sample of phosphorus (red) and bismuth (solid). Point out that as we go from top to bottom in a family (group) of elements on the periodic chart, non-metallic characteristics decrease and metallic characteristics increase. Nitrogen is a gas, phosphorus is a soft solid, and bismuth is a shiny metallic solid.

(C) Show samples of magnesium chloride, calcium chloride, and barium chloride. All are colorless (white) solids.

(D) Show samples of potassium permanganate, copper(II) acetate, copper(II) sulfate, and cobalt(II) nitrate. All are solids that have color. Each contains a transition metal.

(E) Show samples of sugar, napthalene, benzoic acid, and camphor. All of these materials are totally nonmetallic, colorless, and soft.

(F) Show samples of water, ethyl alcohol, and acetone. All of these compounds are liquids and contain only nonmetallic elements.

Note: All samples should be put in containers that may be stored easily and reused for display.

Hazards: CHLORINE GAS IS HAZARDOUS AND SHOULD NOT BE BREATHED. UNLESS YOU CAN WORK WHERE THERE IS ACTUALLY A FLOW OF AIR, DO NOT GENERATE CHLORINE. Just omit that part of this demonstration.

RED PHOSPHORUS IS A DANGEROUS SUBSTANCE (WHITE PHOSPHORUS IS MORE DANGEROUS!) Unless you already have red phosphorus in your stockroom, it is suggested that you do not order it for use in this demonstration. Instead, simply omit phosphorous from this demonstration.

## DEMONSTRATION 8

## PREPARATION OF OXYGEN AND HYDROGEN

### GENERAL TOPIC: VII. DESCRIPTIVE CHEMISTRY

Materials:

1 M aqueous solution of hydrochloric acid, HCl

3% aqueous solution of hydrogen peroxide, $H_2O_2$, (fresh)

magnesium ribbon

active dry yeast (available in ¼ oz. packages at grocery stores)

balance, 0.01 g sensitivity

125 mL Erlenmeyer flask

matches

wood splints

15 x 125 mm test tube

Procedure:

Preparation of Oxygen

Place 30 mL of the 3% $H_2O_2$ solution in the Erlenmeyer flask and add a very small amount of active dry yeast. You should see rapid production of gas bubbles. Ignite a wood splint. Shake out the flame and insert the glowing splint into the neck of the flask. The splint should reignite.

Preparation of Hydrogen

Obtain a 6 cm strip of magnesium ribbon. Roll it loosely into a coil and place in the 15 x 125 mm test tube. Appoint an assistant who, when instructed, will light a match and, place it at the mouth of the test tube. Add enough 1 M HCl to cover the magnesium ribbon. Vigorous gas evolution should occur. Place a cork loosely in the mouth of the test tube. After about 30 seconds request your assistant to place the lighted match at the mouth of the test tube as you remove the cork. A "pop" should be heard.

Notes:

1. This preparation of oxygen gas is recommended over the decomposition of $HgO(s)$ or $KClO_3(s)$. These reactions are too dangerous to be used. The yeast causes the decomposition of hydrogen peroxide to produce oxygen gas and water:

$$2\ H_2O_2(aq) \dashrightarrow 2\ H_2O(l) + O_2(g)$$

$O_2(g)$ supports combustion; thus the glowing spint reignites.

Demonstration 8
Preparation of Oxygen and Hydrogen, page 2

2. The equation for preparation of $H_2(g)$ is

$$Mg(s) + 2\ HCl(aq) \dashrightarrow H_2(g) + Mg^{+2}(aq) + 2\ Cl^-(aq)$$

The "pop" occurs because you ignite a mixture of $H_2(g)$ and $O_2(g)$. This gas mixture is explosive. The equation for this reaction is:

$$2\ H_2(g) + O_2(g) \dashrightarrow 2\ H_2O(l)$$

If you allow the $H_2(g)$ to sweep all of the air from the test tube, the "pop" will not occur and the $H_2(g)$ will only burn because no $O_2(g)$ is left in the tube.

Waste Disposal: Solutions may be discarded by flushing down the sink drain with a large quantity of running water.

## DEMONSTRATION 9

## IGNITION OF METHANE GAS BUBBLES

### GENERAL TOPIC: VIII. ENERGY AND CHEMICAL CHANGE

Materials:

- bubble blowing solution (from discount stores)
- glycerine (glycerol)
- 600 mL beaker
- candle
- 50 mm (top diameter) long-stem 60° glass funnel
- matches
- natural gas outlet with burner fitting
- stick, at least 3 ft. long
- 4 ft of rubber tubing

Procedure: Mix 4 parts of the bubble blowing (soap) solution with one part of glycerine to get larger, longer-lasting bubbles. Stir well. Place this solution in the 600 mL beaker. Connect the funnel to the rubber tubing and attach the other end of the tubing to the natural gas outlet. Place the wide end of the funnel in the soap mixture and adjust the gas pressure to produce fairly large bubbles. Detach each bubble by shaking the funnel. As the bubble rises, you can ignite it with a lighted candle attached to a stick at least 3 ft. long. DO NOT IGNITE THE BUBBLE UNTIL IT IS SAFELY AWAY FROM THE FUNNEL. KEEP FLAME AWAY FROM THE BEAKER AND FUNNEL AT ALL TIMES. DON'T FORGET TO SHUT OFF THE GAS WHEN FINISHED!

Observations: The bubble will rise because the molecular mass of methane is less that that of molecules in air. When ignited the methane burns and the flames expand outward. Energy is released in a dramatic fashion.

$$CH_4(g) + 2\ O_2(g) \longrightarrow CO_2(g) + 2\ H_2O(g)$$

## DEMONSTRATION 10

## EXOTHERMIC AND ENDOTHERMIC PROCESSES

### GENERAL TOPIC: VIII. ENERGY AND CHEMICAL CHANGE

**Materials:**

ammonium nitrate, $NH_4NO_3$
sodium bromide, NaBr
sodium hydroxide, NaOH, pellets
tap water
labels
4 stoppers to fit test tubes
4 large test tubes, Pyrex

**Comments:** When a small amount of sodium hydroxide is dissolved in water, the process is exothermic. When a small amount of ammonium nitrate is dissolved in water the process is endothermic. It should be noted that these processes are not chemical reactions as we usually think of them. A solid substance is simply being dissolved in water. It may have a positive, negative or zero heat of solution. The heat of solution of NaOH is positive, while that of $NH_4NO_3$ is negative. Sodium bromide has a heat of solution of almost zero. Therefore, there is virtually no change in temperature when NaBr is dissolved in water. The process is neither exothermic nor endothermic.

**Procedure:** Pour some sodium hydroxide pellets into a large labeled Pyrex test tube to a height of about 2.5 cm. Into another large labeled test tube pour some solid ammonium nitrate also to a height of about 2.5 cm. Pour into a third labeled Pyrex test tube some sodium bromide to a height of about 2.5 cm. Fill the three test tubes about 3/4 full with tap water. Stopper the test tubes tightly and carefully shake them to dissolve the solids. Finally, fill a fourth labeled Pyrex test tube 3/4 full of tap water. Now place all the tubes in an upright position in a suitable container (such as a small box or a can) and pass the container around the classroom. The students can feel the warmth of the tube containing the NaOH solution and the coolness of the tube containing the $NH_4NO_3$ solution, indicating the exothermic and endothermic processes which occurred. The tube containing the NaBr solution will be at about the same temperature as the tube of tap water used as a reference.

**Hazards:** THE NaOH SOLUTION WILL BE VERY HOT AND VERY CAUSTIC. GREAT CARE IS NEEDED BY THE STUDENTS IN HANDLING THE TEST TUBES TO PREVENT SPILLAGE OR BURNING. IT IS RECOMMENDED THAT THE TEST TUBES BE HANDLED ONLY BY THE VERY TOP OF THE TUBE.

Demonstration 10
Exothermic and Endothermic Processes, page 2

| | |
|---|---|
| Waste Disposal: | When the demonstration is complete, the contents of the test tubes can be carefully emptied in the sink and flushed with large quantities of water. |
| Question: | What are exothermic and endothermic processes? |
| Answer: | Exothermic processes supply heat to their surroundings as they proceed; whereas endothermic processes take heat from their surroundings as they proceed. |

## DEMONSTRATION 11

## A COLLAPSING CAN

### GENERAL TOPIC: IX. GASES AND GAS LAWS

Materials:

Bunsen burner or propane torch

metal can (one gallon size, or smaller) with a cap that will seal the can

some type of graduated cylinder to measure the volume of water

heavy gloves or a pot holder

metal tripod stand or large ring and ringstand.

Procedure:

Add about 20 mL of water to a metal can and place it on a tripod or ringstand. With the cap left off the can, heat the water in the can to boiling with a Bunsen burner or propane torch. After steam has escaped from the open top for about one minute, remove the flame and screw the cap on the can. USE HEAVY GLOVES OR A POT HOLDER WHILE HANDLING THE HOT CAN TO AVOID BURNING YOUR HANDS. Cool the hot can. This may be accomplished by using water from a hose or a wash bottle, cool water in a tray, or a bed or ice.

To Be Emphasized:

This exercise demonstrates the differences in pressures on the outside and inside of the can when the inside pressure is decreased. This is accomplished by first replacing the air inside the can with steam. At 100°C, with the cap off the can, the vapor pressure of water (as steam) inside the can is equal to the atmospheric pressure outside the can. When cool water is poured over the can, with the cap screwed on, the vapor pressure of water inside the can changes from one atmosphere at 100°C to about 1/30 atmosphere at room temperature. The greater outside pressure causes the can to collapse.

Questions:

1. Why does air cause pressure on the can?

   Answer: Moving molecules bump into the can.

2. What would happen if the can were heated with the cap screwed on? (DO NOT ATTEMPT THIS!)

   Answer: The pressure inside the can could increase beyond the limits of the can to withstand the pressure. It would burst, showering the area with scalding water.

## DEMONSTRATION 12

## RELATIVE RATES OF DIFFUSION OF GASES

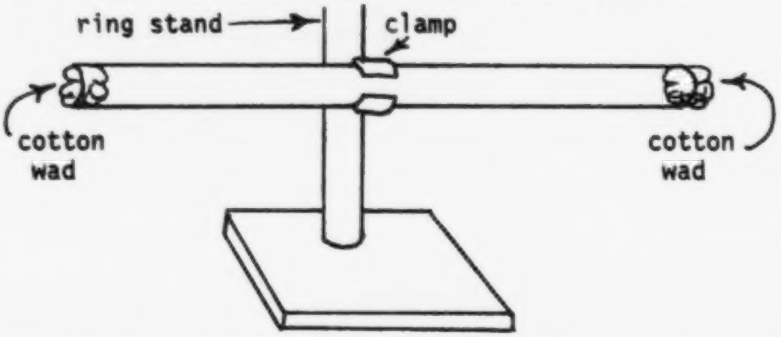
## GENERAL TOPIC: X. KINETIC MOLECULAR THEORY

**Materials:**

- concentrated ammonium hydroxide, $NH_4OH$
- concentrated hydrochloric acid, HCl
- 2 small beakers
- clamp
- cotton balls
- 10 mL graduated cylinder
- gloves or forceps
- meter stick
- ringstand
- glass or clear plastic tube, 13 mm (1/2 in) or larger diameter and 30-100 cm long

**Procedure:**

1. Clamp a tube as shown in the drawing below.



2. Prepare for later use two wads of cotton that will fit firmly into the ends of the tube.

3. Pour about 5 mL of concentrated $NH_4OH$ into one small beaker and about 5 mL of concentrated HCl into a second beaker located some distance away from the first beaker. Avoid breathing fumes of $NH_4OH$ or HCl. Use a hood if available.

4. Using gloves (or forceps) partially dip an edge of one of the cotton wads into $NH_4OH$ and the other into HCl and place the wet edges at opposite ends of the tube at the same instant. Do not allow any of the $NH_4OH$ or HCl solutions to come in contact with your skin or clothing. Rinse the excess HCl and $NH_4OH$ solutions down the drain with plenty of water.

Demonstration 12
Relative Rates of Diffusion of Gases. page 2

5. Allow the system to stand until HCl gas from the HCl solution and $NH_3$ gas from the $NH_4OH$ solution come into contact producing a white cloud of $NH_4Cl$ which appears as a ring within the tube (2-8 minutes).

6. Measure the distances from the wet cotton to the ring.

To Be Emphasized: The demonstration not only demonstrates molecular motion but shows that molecules with a smaller mass move faster than those with a larger mass. It also shows the reaction of two gaseous substances to form a solid product.

Questions:

1. What are the molecular masses of HCl and $NH_3$?

   Answer: 35.5, 17

2. Why does $NH_3$ gas move faster than HCl inside the tube?

   Answer: Molecules with a smaller mass move faster.

3. Even though the molecules move about 1000 miles per hour, why does the gas move so slowly through the tube?

   Answer: The molecules bump into other molecules and change direction millions of times per second.

4. Why is a cloud formed where the $NH_3$ and HCl meet?

   Answer: Solid $NH_4Cl$ is formed.

## DEMONSTRATION 13

## FLAME TESTS

### GENERAL TOPIC: XI. ATOMIC THEORY

Materials:

barium chloride, $BaCl_2$
calcium chloride, $CaCl_2$
copper(II) sulfate, $CuSO_4$ (cupric sulfate)
lithium carbonate, $Li_2CO_3$
1 M aqueous solution of nitric acid, $HNO_3$
potassium chloride, KCl
sodium chloride, NaCl
distilled or deionized water
Bunsen burner
nichrome wire or atomizer

Procedure:

1. The preparation of solutions of small quantities of the salts need not be quantitative. About 0.50 grams (a pinch) of salt dissolved in an estimated 1 to 2 mL quantity of water will be adequate.

2. The Bunsen burner should be adjusted to give a blue flame without making a roaring noise. Prepare a nichrome (chromel-A) wire about 15-20 cm in length with a small loop at the end. Clean the loop end by dipping it into the $HNO_3$ solution and then into the burner flame. This is done until there is little color imparted to the flame by the presence of the wire loop.

3. Dip the wire into the salt solution and then into the flame. Note the color of the flame. This is the characteristic colored flame for the metal ion in the solution. Repeat the procedure if necessary to get good color.

4. Rinse the wire in water and clean by dipping it into the $HNO_3$ solution and heating between uses in the test solutions.

5. After the students have observed flame tests using the known solutions, ask the class to identify an unknown solution.

6. A grating or prism spectroscope will show the atomic line spectra.

Demonstration 13
Flame Test, page 2

Note: An atomizer can be used instead of a nichrome wire to introduce the solutions to the flame; however, the likelihood of contamination of solutions is greater with an atomizer.

To Be Emphasized:

1. The flame test is useful in determining the identity of many metals in ionic compounds of that metal.
2. The colors demonstrated are due to the metal's atomic emission of light. The flame is energetic enough to dissociate the compound into metal atoms in the flame.
3. The characteristic "line" spectra can be viewed through a simple spectroscope.
4. The colors displayed by a selection of salts:

| Salt | Color Due To | Color |
|---|---|---|
| NaCl | Na atoms | yellow |
| $CaCl_2$ | Ca atoms | yellow red |
| $Li_2CO_3$ | Li | carmine |
| $CuSO_4$ | Cu | emerald green |
| $BaCl_2$ | Ba | yellowish green |
| KCl | K | violet |

5. Flame tests are used to measure sodium and potassium levels in blood.

Waste Disposal: With one exception, all solutions may be flushed directly down the sink drain with running water. A water solution of some soluble sulfate salt or a dilute solution of sulfuric acid should be added to the barium chloride solution to precipitate the very insoluble barium sulfate which is harmless. The mixture of solutions along with the barium sulfate may be washed down the drain with running water.

## DEMONSTRATION 14

## SOLUBILITY IN POLAR AND NONPOLAR SOLVENTS

### GENERAL TOPIC: XIII. SOLUTIONS

Materials:

chalk (calcium carbonate, $CaCO_3$)
ethanol or methanol
kerosene
naphthalene
sodium chloride, NaCl
sugar
10 mL graduated cylinder
12 glass stirring rods
12 test tubes

Procedure: Test the solubility in water of the compounds named in the following list by adding small amounts of the compounds to about 5 mL of water in test tubes: kerosene, water, sugar, naphthalene, chalk, and ethanol. Observe and note whether the compounds dissolve on stirring. Repeat, using kerosene instead of water as a potential solvent.

To Be Emphasized: If a compound dissolves in water (polar) and not in kerosene (nonpolar), the compound is most likely ionic or polar covalent. Polar solvents dissolve ionic compounds if they can be dissolved; nonpolar solvents will not dissolve ionic compounds. If a compound dissolves in kerosene and not in water, it is most likely nonpolar covalent.

Questions: What is suggested concerning the molecular or ionic nature of a compound when it:

1. dissolves in water but does not dissolve in kerosene?

   Answer: Compound is either polar covalent or ionic.

2. dissolves in kerosene but does not dissolve in water?

   Answer: Compound is neither polar covalent nor ionic.

3. dissolves in both kerosene and water?

   Answer: Part of the molecule is like water; another part is like kerosene.

4. does not dissolve in water or kerosene?

   Answer: In the case of this particular substance, solubility is not a useful property for determining the type of compound.

Demonstration 14
Solubility in Polar and Nonpolar Solvents, page 2

Waste Disposal:

A. Of substances involved in the water solubility tests.
Water solutions of sugar and ethanol may be washed down the sink.

Dispose of the chalk-water mixture, the kerosene-water mixture and the naphthalene-water mixture by the procedure described below for kerosene test mixtures.

B. Of substances involved in the kerosene solubility tests.
Put some crumpled paper towels in a small plastic bag. Pour all of the kerosene test mixtures on the towels so that they will be absorbed. Tightly seal the plastic bag and discard it in a trash container.

## DEMONSTRATION 15

## IMMISCIBLE LIQUIDS

### GENERAL TOPIC: XIII. SOLUTIONS

Materials:

food color

250 mL glycerin (sometimes spelled "glycerine" or called glycerol)

250 mL mineral oil

narrow, flat, transparent bottle with top (approximately 700 mL capacity)

Procedure:

Pour the glycerin into the bottle. Add several drops of food color and mix to form a uniform color. Add the mineral oil to the bottle. Cap the bottle and gently shake the contents . Notice the wave-like motion of the interface of the two liquids. Mineral oil is non-polar and glycerin is polar, therefore, the two liquids will not mix. Food colors are polar so they will dissolve in the glycerin but not in the mineral oil. Since the mineral oil is much less dense than the glycerin, it will float on the top of the glycerin. The difference in densities and the fact that the two liquids are immiscible causes the wave-like fluctuation of the interface when the bottle is shaken.

Waste Disposal:

Glycerin is used in food, cosmetics, and pharmaceuticals and is safe to flush down the drain. Mineral oil is used as a laxative and in skin care preparations and is safe to flush down the drain.

Questions:

1. What does immiscible mean?

   Answer: Incapable of mixing

2. Should mineral oil and water mix?

   Answer: No, because mineral oil is nonpolar and water is polar.

3. Why is the glycerin the bottom layer in the experiment?

   Answer: The glycerin layer is on the bottom because it has the greater density.

## DEMONSTRATION 16

## DILUTION OF SOLUTIONS

### GENERAL TOPIC: XIII. SOLUTIONS

**Materials:**

- potassium permanganate, $KMnO_4$
- 10% aqueous solution of sodium hydrogen sulfite, $NaHSO_3$ (sodium bisulfite)
- sulfuric acid, $H_2SO_4$ (either concentrated or dilute)
- balance, 0.01 g sensitivity
- four 250-500 mL graduated beakers,
- 25, 50 or 100 mL graduated cylinder
- pH paper

**Procedure:**

1. Weigh 2-3 g of potassium permanganate into graduated beaker #1. Fill with water to the 250 mL mark and stir until the solid dissolves.
2. Transfer 25 mL of solution #1 to graduated beaker #2, fill with water to the 250 mL mark, and stir.
3. Transfer 25 mL of solution #2 to graduated beaker #3, fill with water to the 250 mL mark, and stir.
4. Transfer 25 mL of solution #3 to graduated beaker #4, fill with water to the 250 mL mark, and stir.

**To Be Emphasized:**

1. This exercise shows how a concentrated solution may be diluted to very dilute concentrations without measuring unusually large or small volumes.
2. The exercise shows the relative color intensities of solutions over a wide range of concentrations.

Dilution of Solutions, page 2
Demonstration 16

Question: What are the molar concentrations of the four solutions?

$$\#1 \quad \text{Molarity} = \frac{\text{g } KMnO_4}{0.250 \text{ L soln.}} \times \frac{1 \text{ mole } KMnO_4}{158 \text{ g } KMnO_4}$$

$$\#2 \quad \text{Molarity} = \text{Molarity } \#1 \times \frac{25 \text{ mL}}{250 \text{ mL}}$$

$$\#3 \quad \text{Molarity} = \text{Molarity } \#2 \times \frac{25 \text{ mL}}{250 \text{ mL}}$$

$$\#4 \quad \text{Molarity} = \text{Molarity } \#3 \times \frac{25 \text{ mL}}{250 \text{ mL}}$$

Waste Disposal: Permanganate ions, $MnO_4^-$, should be reduced to $Mn^{+2}$ ions for disposal. Add sulfuric acid (either concentrated or dilute) CAREFULLY with stirring to each beaker of permanganate solution until the pH of the solution is less than 3.0 (as indicated by pH paper). Only a few drops of sulfuric acid should be required for each beaker. Then add a 10% aqueous solution of sodium bisulfite with stirring until the purple color of the permanganate disappears. About 50-60 mL of $NaHSO_3$ solution will be required to decolorize solution #1; approximately 6 mL to decolorize solution #2; about 3 mL for solution #3; and approximately 2 mL for solution #4. The contents of the beakers may now be flushed down the sink drain with a large quantity of running water.

# DEMONSTRATION 17

## SIMULATED CHEMICAL EQUILIBRIUM

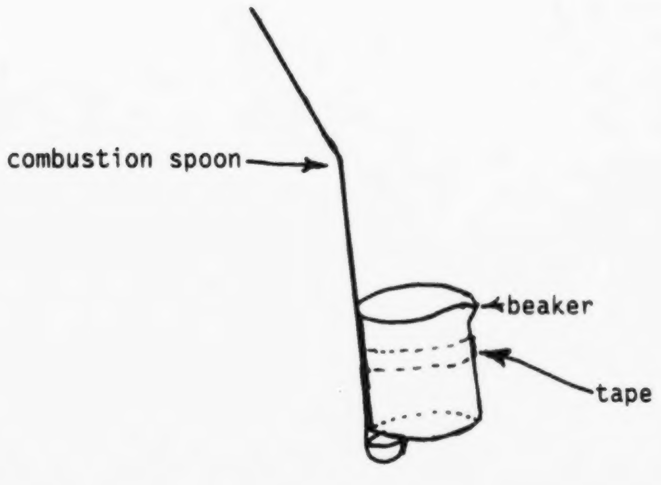
## GENERAL TOPIC: XIV. EQUILIBRIUM

Materials:

- food coloring
- aprons
- 100 mL beaker
- 400 mL beaker
- two 3 L semi-transparent plastic buckets
- two large labels (to identify student helpers)
- combustion spoon

Procedure: Fill one of the plastic buckets 3/4 full with water to which some food coloring has been added. Use two student volunteers. One student is labeled reactants and is given the 400 mL beaker. The other student is labeled products and uses the 100 mL beaker attached by tape to the lower part of the combustion spoon (see diagram below). Then reactants uses his/her beaker to transfer 400 mL of water from the bucket of water to the second bucket which is empty at the start. Products then uses his/her beaker to transfer water from the product (second) bucket back to the reactant (first) bucket. The students continue to alternate transfers for at least 20 minutes.

Observations: After at least 20 minutes of transfer you should notice that the two liquid levels remain nearly constant even as the transfer continues. This is an example of a dynamic state of equilibrium which is typical of chemical systems.



## DEMONSTRATION 18

## pH OF COMMON SUBSTANCES

### GENERAL TOPIC: XV. ACIDS AND BASES

Materials:

antacid tablets (wetted)
"Drano" or other solid drain cleaner (wetted)
various fruits (such as grapefruit, lemons, apples)
"Liquid-Plumr" (liquid drain cleaner)
saliva
vinegar
tap water
small containers (glass or plastic)
safety glasses
gloves
paper towels
pH paper
glass stirring rod

Procedure:

Arrange a display of the common substances to be used. Use pH paper to test the pH of the items displayed.

Put a small amount of solid or liquid to be tested into a labeled container. If testing a liquid, dip the stirring rod tip into the liquid in the small container and touch the pH paper with the wet stirring rod. This is better than dipping the pH paper into liquids. Remember to rinse the stirring rod under running water and wipe it after testing each substance.

This demonstration shows that few common substances are exactly pH 7 or neutral.

Citrus fruits are acidic; antacid tablets are basic. Drain cleaners are strongly basic since most have a high concentration of NaOH.

Saliva can range from below pH 7 to above pH 7.

Hazards:

Only the two drain cleaners present a hazard. USE SAFETY GLASSES AND GLOVES WHEN TESTING THESE CHEMICALS. Dispose of them by flushing them down the sink drain with plenty of water.

Demonstration 18
pH of Common Substances, page 2

Questions:

1. Some plants grow best in acidic soil; others grow best in basic to neutral soil. How would you test the pH of soil?

   Answer: Put a small amount of the soil to be tested in a test tube, fill the tube about 1/2 to 2/3 full of water, shake vigorously, let the solid matter settle, and test the solution with pH paper.

2. Gastric juices in the stomach are approximately pH 2. Are these juices strongly acidic, weakly acidic, or neutral?

   Answer: strongly acidic

Note: For maximum effectiveness of this demonstration, it should be presented to a small group of people so that all can see clearly the changes in the pH paper.

## DEMONSTRATION 19

## ACIDS AND PENNIES

### GENERAL TOPIC: XVI. OXIDATION-REDUCTION

**Materials:**

50 mL of 6 M aqueous solution of hydrochloric acid, HCl
50 mL of 6 M aqueous solution of nitric acid, $HNO_3$
2 pre-1980 pennies (100% copper)
distilled or deionized water
tap water
apron
2 100 mL beakers
damp cloth
100 mL graduated cylinder
2 500 mL Erlenmeyer flasks
gloves

**Preparation of Solutions:**

6 M HCl is prepared by adding a given volume of concentrated HCl (which is 12 M) slowly with stirring (CAREFUL!) into 25 mL of distilled or deionized water in a 100 mL beaker.

$HNO_3$ is prepared by pouring 19 mL of concentrated nitric acid (1.42 specific gravity; 15.8 M) slowly with stirring (DANGER! BE CAREFUL!) into 31 mL of distilled or deionized water in a 100 mL beaker.

**Procedure:**

1. Place a pre-1980 penny in the bottom of each Erlenmeyer flask.
2. Pour 50 mL of 6 M $HNO_3$ into one flask and 50 mL of HCl into the other flask.
3. Observe brown-red fumes in the flask with $HNO_3$. There is no activity in the flask with HCl. Point out that both HCl and $HNO_3$ are acids. The reaction with $HNO_3$ must be caused by something other than action of an acid. The copper penny is being oxidized to copper(II) ions by the $NO_3^-$ ions of the nitric acid. However, it is important to point out that $NO_3^-$ ions are not good oxidizing agents unless the solution is very acidic. The penny would not be oxidized in a water solution of $NaNO_3$. Thus, $H^+$ ($H_3O^+$) ions as well as $NO_3^-$ ions must be present for oxidation.

Demonstration 19
Acids and Pennies, page 2

4. Before the very toxic brown-red $NO_2$ fumes reach the top of the flask, quench the reaction by adding about 150 mL of tap water. Cover the flask with a damp cloth to prevent the $NO_2$ fumes from escaping into the room.

Hazards: BOTH ACID SOLUTIONS ARE CORROSIVE. EYE PROTECTION IS ESSENTIAL. YOU MAY ALSO WANT TO WEAR A LABORATORY APRON OR COAT AND POSSIBLY GLOVES.

Questions:

1. Are the $NO_3^-$ ions of nitric acid oxidized or reduced?

   Answer: reduced

2. What is the oxidizing agent, copper or nitric acid?

   Answer: $HNO_3$

3. In a chemical reaction can you have oxidation without a corresponding reduction?

   Answer: no

4. Does the copper lose or gain electrons when it dissolves?

   Answer: loses electrons -- This is oxidation.

Waste Disposal: At the conclusion of the demonstration dispose of the solutions by filling the Erlenmeyer flasks with water and pouring the diluted acids down a sink drain with a large amount of running water.

## DEMONSTRATION 20

## MEASUREMENT OF THE pH OF WATER SAMPLES

### GENERAL TOPIC: XIX. ENVIRONMENTAL CHEMISTRY

**Materials:** glass or plastic containers with caps

pH paper (both wide- and narrow-range)

**Procedure:** Collect about ten 100 mL water samples in clean glass or plastic containers and secure the cap on the containers. These may include: tap water, rain water, bath water after taking a bath, dish water, a small amount of tap water that has been boiled a few minutes and then cooled without stirring, industrial waste water, well water, pond water, 100 mL of water that has about 20 g of soil stirred in it and allowed to stand a few minutes, and sea water. Test the pH of each sample by dipping small lengths of first wide-range and then narrow-range pH paper into the container.

**Questions:**

1. What is the pH of a neutral solution?

   Answer: pH = 7

2. What is the pH of bath or dish water?

   Answer: greater than 7

3. If tap water changes pH on boiling, what accounts for the change?

   Answer: removal of dissolved $CO_2$

4. What is commonly added to soil to increase its pH?

   Answer: lime, CaO

# CHAPTER 8

# EXPERIMENTS

## CHAPTER 8 -- EXPERIMENTS

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# EXPERIMENT 1

# THE BURNING OF A CANDLE

## GENERAL TOPIC: I. INTRODUCTION

| | |
|---|---|
| Introduction: | What does a candle flame really look like? Do we know? We will use our powers of observation for this experiment. |
| Objective: | To learn to make careful observations. |
| Materials and Equipment: | 400 mL beaker<br>1000 mL beaker<br>Bunsen burner<br>candle<br>cardboard<br>cotton ball (100 count)<br>500 mL Erlenmeyer flask<br>forceps<br>watch glass<br>wire gauze<br>matches<br>plain white paper<br>timer or watch. |
| Chemicals: | 1 lb. sugar (cubes)<br>lampblack |
| Procedure: | Prepare a table for recording data as shown below. For each section of the experiment, record the results obtained for that number. See the example. |

| Section | Observation |
|---|---|
| A.1 | A. 1 The flame rises, color of flame was..... |

A. The Nature of the Flame:

1. Set the candle on a piece of cardboard and let it burn for a minute or two. Observe the flame.

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The Burning of a Candle, page 2

2. Hold the cotton ball with forceps and light it. Compare the flame with the candle flame.

3. Observe the liquid wax just below the flame of the candle.

4. Turn the candle upside down on a piece of paper and note what happens.

5. Place about 1 mL of water on a watch glass; add a lump of sugar. Describe what happens to the water and the sugar. How is this action similar to that involved with a candle wick?

6. Snuff out the flame of the candle and note what happens.

7. Relight the candle, snuff it out, and then quickly bring a lighted match within an inch or two of the wick. Observe.

B. The Structure of a Candle Flame:

1. Place a sheet of white paper behind the flame. Look through the portion of the flame surrounding the wick.

2. Hold the sheet of paper horizontally and lower it into the flame momentarily so that the paper touches the candle wick, pulling the paper away just before it catches fire. Observe where the paper is most burned. Examine the side of the paper next to the flame.

3. Place a pinch of lampblack on a sheet of paper. Roll paper into a tube and blow lampblack into the flame of a Bunsen burner. Note the result.

4. Bring the edge of a wire gauze into the candle flame. Note the result.

C. A Requirement for Combustion:

1. Invert a perfectly clean 400 mL beaker over the burning candle. Note the result.

2. After wiping the beaker out thoroughly and laying it on its side for a few minutes, repeat the experiment and record the time needed for the flame to be extinguished. Repeat two times. Find the average of the three periods.

3. Repeat step 2 with a clean 1000 mL beaker.

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The Burning of a Candle, page 3

4. Invert a 400 mL beaker over the burning candle, letting the flame go out. Quickly remove the beaker and set it down on the desk top, still inverted. Relight the candle and quickly place the beaker over it. Note the result.

5. Wipe the beaker out completely. Relight the candle. Take a deep breath, hold for 30 seconds, then exhale slowly into the beaker. Quickly invert the beaker over the flame. Note the result.

6. Hold a cold metal object above the candle but not in a position to be sooted. Note what happens.

Questions:

1. What is the purpose of the wick in a candle?

2. Will a candle of larger diameter produce a larger flame?

# EXPERIMENT 1 -- TEACHER'S GUIDE

# THE BURNING OF A CANDLE

## GENERAL TOPIC: I. INTRODUCTION

Pre-laboratory Discussion: The difference between interpretation and observation should be discussed with the students. This may be the student's first lab experiment and they will come into the lab with varied experiences. A demonstration to show that observations are made by using the senses would be helpful. For example, you might have a brown paper bag with a few oranges in it. Allow the students to ask pertinent questions. Encourage them to smell the contents, feel through the paper bag, and attempt to identify the unknown.

Sample Data: A. THE NATURE OF THE FLAME

1. The flame rises because heat causes the gases to expand and the density of the gases to decrease.
2. The cotton ball burned with a yellow flame and was extinguished fairly rapidly. The flame of the candle was steady and gave off more light than the cotton ball.
3. There is liquid wax in the cup at the top of the candle. This wax does not burn because only vaporized wax burns well.
4. When the candle was turned upside down, the flame did not go completely out. A small flame continued to burn. The heat of the candle serves two purposes: (a) to melt and vaporize the wax and (b) to provide the necessary kindling temperature.
5. The water was drawn upward into the sugar cube through capillary action. The cube then dissolved. Two other examples of capillary action are paper towels absorbing liquids and water rising through plant stems.
6. When the flame was pinched out smoke rose from the candle. There was a smell of vaporized wax.
7. When the flame was snuffed out and a lighted match brought near the wick, the candle relighted. This was done quickly after snuffing out the candle so the wax would still be vaporized.

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## B. THE STRUCTURE OF A CANDLE FLAME

1. The transparent part of the flame surrounding the wick was seen by placing the paper behind the flame.

2. Holding a piece of white paper above the flame produced a dark ring on the bottom side of the paper. The area around the outer edges was slightly darker than the center. This occurred because the hottest part of a flame is near the edge where there is direct contact with air. No soot was found on the bottom of the paper.

3. Soot is obtained from the bottom of a smoked watch glass. Blowing the lampblack into the flame of a bunsen burner produced yellow bursts when the lampblack hit the flame.

4. The flame burned below the gauze but not above it because the metal absorbs so much heat that the kindling temperature is not reached above the gauze. Bringing a lighted match above the gauze near the edge of the flame caused the flame to be rekindled above the gauze.

## C. A REQUIREMENT FOR COMBUSTION

1. When a 400 mL beaker was inverted over the flame, the flame was extinguished because of lack of air.

2. Time required for the flame to be extinguished:

| | | |
|---|---|---|
| 400 mL beaker | Trial 1 | 14.8 sec |
| | Trial 2 | 15.2 sec |
| | Trial 3 | 13.1 sec |
| | Average | 14.4 sec |

3. Time required for the flame to be extinguished:

| | | |
|---|---|---|
| 1000 mL beaker | Trial 1 | 29.5 sec |
| | Trial 2 | 27.3 sec |
| | Trial 3 | 28.7 sec |
| | Average | 28.5 sec |

This experiment showed that air (oxygen) is required for combustion and that the time required for the flame to be extinguished is proportional to the amount of air present. Some students may make the observation that there was condensed vapor on the sides of the beaker.

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4. When the beaker was placed over the second flame immediately after the first flame had been extinguished, and without replenishing the air in the beaker, the second flame was extinguished rapidly.

5. The time required for the flame to be extinguished with the beaker containing exhaled air was 6.7 seconds.

6. There was vapor on the bottom of a cold metal spoon. The vapor was more noticeable on the sides of the beaker than on the spoon.

Answers to Questions:

1. The purpose of the wick is to allow a place for the melted wax to burn. The melted wax flows up the wick.

2. A candle of larger diameter would not produce a larger flame unless the wick also is larger. Multiple wicks may be used. After the students have answered this question a demonstration using a larger candle and multiple wicks would be appropriate.

Post-laboratory Discussion: This laboratory experiment lends itself to an active class discussion while the teacher lists on the chalkboard or the overhead projector the varied responses of the students. Answers will vary greatly.

## EXPERIMENT 2

## MASS AND VOLUME RELATIONSHIP

### GENERAL TOPIC: II. METRIC SYSTEM AND DENSITY

Introduction: Mass and volume are different properties of an object. Mass is the amount of matter in an object. It is measured by comparison with a known mass using a balance. The volume of an object is the space occupied by the object. The volume of a solid object can be determined by measuring the volume of water displaced by the object.

One of the important jobs that chemists do is to discover the composition of an object. They can do this by examining the properties of the object. In this experiment we will explore whether the mass and volume of an object can be used to identify the substance of which the object is made.

Objective: The purpose of the experiment is to investigate the properties of mass and volume of some objects and to discover how mass and volume can be used to identify the material of which an object is made.

Materials and Equipment:

balance, 0.01 g sensitivity

50 or 100 mL graduated cylinder

thread

Chemicals:

piece of aluminum metal

piece of brass metal

piece of unknown metal

Procedure:

1. Weigh the brass object and the aluminum object on the balance. Record their masses in Table 1.

2. Determine the volume of the brass object and the aluminum object by displacement of water as demonstrated by your teacher. Record your data on lines 1 and 2 of Table 2. After calculating the volume of the object on line 3 of Table 2, enter the values of the volumes of the objects on line 2 of Table 1.

3. Determine the mass/volume ratios for aluminum and the brass objects. Record these in Table 1.

4. You have been given an "unknown" object which is made of either brass or aluminum. See if you can figure out what the unknown object is made of by comparing its properties with those of the "known" objects. As you do this part

Experiment 2
Mass and Volume Relationship, page 2

of the experiment enter your data under "unknown object" in Tables 1 and 2. Begin by comparing such properties as color and shape. Then compare masses, volumes. and mass/volume ratios. Answer the questions in the questions section of the experiment as you do this comparison. Be sure to record the unknown object number in the blank in Table #1.

Data:

Table 1

| | Aluminum object | Brass object | Unknown object #____ |
|---|---|---|---|
| Mass (g) | | | |
| Volume (mL) | | | |
| Mass/Volume Ratio | | | |

Table 2

| | Aluminum object | Brass object | Unknown object #____ |
|---|---|---|---|
| Volume of water (mL) | | | |
| Volume of water and object (mL) | | | |
| Volume of object (mL) | | | |

Questions:

1. Can you identify the unknown metal by examining it? Explain.

2. Can you identify the unknown metal by comparing its mass with the masses of the known objects? Explain .

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Mass and Volume Relationship, page 3

3. Can you identify the unknown metal by comparing its volume with the volumes of the known metals?

4. Now try to identify the unknown metal by comparing the ratio of its mass to its volume (mass/volume) with the mass to volume ratio of the known metals. Show your work and write your conclusion in this space.

Further Study:

1. If you had a piece of aluminum with a volume of 13 mL, what would its mass be? Show your work.

2. What would be the mass of a piece of brass which has a volume of 65 mL? Show your work.

3. If a piece of aluminum had a mass of 9 grams what would its volume be? Show your work.

## EXPERIMENT 2 -- TEACHER'S GUIDE

## MASS AND VOLUME RELATIONSHIP

### GENERAL TOPIC: II. METRIC SYSTEM AND DENSITY

**Introduction:** This experiment illustrates the relationship between the properties of mass and volume and the concept of density.

**Objectives:**

1. The student should learn how to measure the volume of an object by water displacement.
2. The student should be able to make comparisons of properties of the objects.
3. To discover the property of density and to see that the density can be used to identify the material of which an object is made.
4. To be able to write clear explanations or answers to questions.

**Pre-laboratory Discussion:** The students should be somewhat familiar with the properties of mass and volume before doing this experiment. Students should know how to use the balance and how to read a graduated cylinder. The teacher should demonstrate how to find the volume of an object by displacement of water before beginning the experiment. (To avoid breaking a glass graduated cylinder, tie a thread around the metal object before lowering it into the graduated cylinder.)

Density is not introduced by defining it and measuring the density of an object as in the standard experiment. Instead, the concept of density is introduced by a discovery method.

After the students obtain the mass and volume of two metal objects, they are asked to explore whether they can identify an unknown metal object by comparing its properties with the known metal objects. Hopefully, they will discover that mass and volume are not properties specific to a certain substance (intrinsic properties) but that samples of a substance can have any mass and/or volume. It is the ratio of mass to volume, or the mass of a fixed volume which is the unique property of a substance.

**Preparation of Materials:** Balances which are graduated in 0.1 g increments can be used.

For each student or pair of students:
1 50 or 100 mL graduated cylinder
2 pieces of "known" metal.

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These should be different metals such as aluminum and brass. They can have the same or different masses and volumes. It is suggested (if possible) that you use cylinders of brass and aluminum or other metals which can be made by having 1/2 inch rod cut into lengths such as 5 cm, 6.2 cm, 7.5 cm.

Cylinders can be obtained from scientific supply houses. These are listed under specific gravity sets.

Another suggestion is to use different size bolts.

One piece of unknown metal. This should be a piece of one of the known metals disguised by painting it with black paint. The piece of unknown metal should not be the same mass or volume as that of the known metal assigned to a student.

Post-laboratory Discussion: Introduce the term density as the mass/volume ratio. Take one particular metal e.g. aluminum and compare the densities of 3 or more different sized pieces using the students' data. You want to illustrate the fact that although the masses and volumes may be different, the densities of the different pieces of a metal are the same. You might also want to compare the masses of a piece of brass and a piece of aluminum having the same volume.

Show students how to use the mass/volume ratio or volume/mass ratio as a conversion factor to solve the problems assigned under "Futher Study". For example:

What is the mass of a piece of lead with a volume of 17 mL?

The density or mass/volume ratio of lead is 11.2 g/mL

$$\text{Mass} = 17\ \text{mL} \times \frac{11.2\ \text{g}}{\text{mL}} = 190\ \text{g}$$

What is the volume of a piece of lead which has a mass of 25 g?

$$\text{Volume} = 25\ \text{g} \times \frac{1\ \text{mL}}{11.2\ \text{g}} = 2.2\ \text{mL}$$

Students may want to compare their value for density with that in a chemistry handbook. Additionally, a discussion of Archimedes Principle might be well received at this time.

# EXPERIMENT 3

## SCIENTIFIC NOTATION AND INDIRECT MEASUREMENTS

### GENERAL TOPIC: III. SCIENTIFIC NOTATION AND SIGNIFICANT FIGURES

Introduction: Scientists must make many measurements. Often these measurements involve very large or very small numbers. It is important that these values represent the accuracy of the measurements. To do this, values are reported in significant figures.

Very large and very small numbers are difficult to express in the ordinary way. To save space and to make these numbers easier to recognize, they are expressed in scientific notation.

Many things can be measured directly. Using a ruler, we can measure the length, width, and thickness of a book. However if the thickness of a sheet of paper is needed, this would be difficult to measure directly. For making such small measurements an indirect method is usually more accurate. The thickness of five hundred sheets can be measured directly and the thickness of one sheet calculated from an indirect measurement.

The following metric relationships are useful in expressing measures:

| | | |
|---|---|---|
| mega (M) | 1,000,000 | $10^6$ |
| kilo (k) | 1,000 | $10^3$ |
| hecta (h) | 100 | $10^2$ |
| deka | 10 | $10^1$ |
| Base | 1 | $10^0$ |
| deci (d) | 0.1 | $10^{-1}$ |
| centi (c) | 0.01 | $10^{-2}$ |
| milli (m) | 0.001 | $10^{-3}$ |
| micro ($\mu$) | 0.000,001 | $10^{-6}$ |
| nano (n) | 0.000,000.001 | $10^{-9}$ |

Objectives:

1. To learn to recognize the accuracy of measurements and report information calculated from them using appropriate significant figures.
2. Express large and small numbers in scientific notation.
3. Make indirect measurement of the thickness of paper and the volume of a drop of water.

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Scientific Notation and Indirect Measurements, page 2

**Materials and Equipment:**

- balance, 0.01 g sensitivity. Also there must be access to a second balance that has the capacity to weigh a ream of paper rather accurately.
- beakers (50 or 100 mL)
- evaporating dish
- medicine dropper
- ream of paper (500 sheets)
- ruler (in metric units)

**Procedure and Calculations:**

## A. Significant Figures

In making a calculation that involves a measured value, all figures that have meaning are used. They are figures about which there is no doubt plus one estimated figure. Significant figures should not be confused with the number of decimal places. Each of the following numbers, 4230, 423, 42.3, 4.23, and 0.00423, has three significant figures. The zeros in the first and last numbers locate the decimal place.

In the numbers 4230. and 0.004230 the zeros on the right are significant and indicate a value with four significant figures. In these numbers, the zeros on the right are not required to locate the decimal place.

When a value is accurate to three significant figures and we have four reported, the last figure is rounded off. If it is greater than five, the third figure is increased by one, 3.257 would round off to 3.26. If the last figure is less than five, the third figure remains as it is. 42.62 would round off to 42.6. If the last figure is five, and the third figure is even, the third figure remains as it is. 5.285 would round to 5.28. If the last figure is five and the third figure is odd, the third figure increases by one. 73.35 would round to 73.4.

A student measures the length of a table and reports it as 90.3 cm. His ruler is marked off such that he can estimate a tenth of a centimeter. A second student measures the width of the table and reports it as 54.48 cm. His ruler is marked off such that he can estimate to a hundredth of a centimeter.

What is the sum of the length and distance?

$$\begin{array}{r} 90.3\ \ \text{cm} \\ \underline{54.48\ \text{cm}} \\ 144.78\ \text{cm} \end{array}$$

This value should be reported as 144.8 cm since the least accurate measure was to a tenth of a centimeter.

The area of the above table is:

$$90.3 \text{ cm} \times 54.48 \text{ cm} = 4919.5 \text{ cm}^2$$

This product should be reported as 4920 cm$^2$. Even though the least accurate measurement was to a tenth of a centimeter, our answer is no longer in centimeters but is in a derived unit (square centimeters). The least accurate measurement contains three significant figures and that is what we must report in our product with its derived unit.

When counted numbers or defined numbers are used in calculations, the degree of accuracy depends upon the accuracy of the measured numbers with which the counted or defined numbers are used. If a group of boxes has approximately 320 apples per box, then 3 boxes would have 3 boxes x 320 apples per box = 960 (not 1000 apples). Two significant figures are in our answer even though the "3" is one figure. The "3" is a counted value and has as much accuracy as we need. When finding the number of feet in 4330 inches,

$$4330 \text{ inches} \times \frac{1 \text{ ft}}{12 \text{ inches}} = 361 \text{ ft (not 360 ft)},$$

three significant figures are in our answer even though the number "12" (of the 12 inches term) has two figures. 12 inches = 1 foot is a defined set of numbers.

Express the following in the proper number of significant figures.

(1) 25.66 + 38.205 = ____________________

(2) 46.22 - 2.2183 = ____________________

(3) 25.66 x 38.235 = ____________________

(4) 46.2 ÷ 2.7843 = ____________________

## B. Scientific Notation

Very large and very small numbers are more conveniently expressed with one digit to the left of the decimal and the value multiplied by ten to the proper exponent.

$$253{,}400{,}000 = 2.534 \times 10^8$$

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Scientific Notation and Indirect Measurements, page 4

The decimal was moved to the left by eight digits, thus we multiply the value by $10^8$.

Consider the number below:

$$0.00002510 = 2.510 \times 10^{-5}$$

The decimal was moved to the right by five digits; thus we multiply the value by $10^{-5}$.

In adding or subtracting numbers expressed in scientific notation, both values must be in the same exponent of ten.

$$4.25 \times 10^{-3} + 2.83 \times 10^{-2} = 0.425 \times 10^{-2} + 2.83 \times 10^{-2}$$
$$= 3.26 \times 10^{-2}$$

In multiplying numbers expressed in scientific notation, we multiply the numerical values and add the exponents.

$$(2.78 \times 10^{-7})(7.6 \times 10^{-3}) = 21 \times 10^{-10} = 2.1 \times 10^{-9}$$

$$(2.78 \times 10^{-7})(7.6 \times 10^{3}) = 21 \times 10^{-4} = 2.1 \times 10^{-3}$$

In dividing numbers expressed in scientific notation, we divide the numerical values and subtract the exponents.

$$(2.78 \times 10^{-7}) \div (7.6 \times 10^{-3}) = 0.37 \times 10^{-4} = 3.7 \times 10^{-5}$$

$$(2.78 \times 10^{-7}) \div (7.6 \times 10^{3}) = 0.37 \times 10^{-10} = 3.7 \times 10^{-11}$$

Scientific notation is useful to indicate significant figures in large numbers when zeros are involved. If the value 25000 is known to three significant figures, it can be expressed as $2.50 \times 10^4$.

Express the following numbers in scientific notation keeping the correct number of significant figures.

(1) $0.00234 =$ ____________
(2) $52300 =$ ____________
(3) $0.000200 =$ ____________
(4) $3.62 \times 10^{-3} - 1.75 \times 10^{-2} =$ ____________
(5) $0.004304 + 2.18 \times 10^{-3} =$ ____________
(6) $2.50 \times 10^{-7} \times 3.11 \times 10^{2} =$ ____________
(7) $(8.1 \times 10^{16}) \div (2.77 \times 10^{4}) =$ ____________
(8) $(6.234 \times 10^{-5}) \div (3.4 \times 10^{3}) =$ ____________

## C. How Can You Determine The Thickness And Mass Of A Single Sheet Of Paper?

Measure the thickness of a ream (500 sheets) of paper and then calculate the thickness of one sheet.

(1) Thickness of 500 sheets __________cm
(2) Thickness of 1 sheet in cm __________cm
(3) Thickness of 1 sheet in μm __________μm

Weigh the whole ream of paper, calculate the mass of a single sheet, and then the mass of 1 $cm^2$ of a single sheet.

(4) Mass of 500 sheets __________g
(5) Mass of 1 sheet __________g
(6) Mass of 1 sheet in mg __________mg
(7) Length of paper __________cm
(8) Width of paper __________cm
(9) Area of 1 sheet of paper __________$cm^2$
(10) Mass of 1 $cm^2$ of 1 sheet of paper in g __________g
(11) Mass of 1 $cm^2$ of 1 sheet of paper in mg __________mg
(12) Mass of 1 $cm^3$ of paper in g __________$g/cm^3$

## D. How Can You Determine The Volume Of One Drop Of Water?

Carefully balance a dry evaporating dish on the balance and then add 3.0 grams of weight to the beam.

Using a medicine dropper, count drops of water into the dish until you have added 3.0 grams.

(1) Number of drops in 3.0 grams of water ____drops
(2) Volume of 3.0 grams of water (Assume density of water is 1.0 g/mL) ____mL
(3) Volume of 1 drop of water (in mL) ____mL
(4) Volume of 1 drop of water (in μL) ____μL
(5) Mass of 1 drop of water (in mg) ____mg
(6) Mass of 1 drop of water (in μg) ____μg

## E. How Many Molecules Are In One Drop Of Water?

The molecular mass of water in grams is called a mole of water and contains $6 \times 10^{23}$ molecules. Using the data from part D above, calculate the number of drops in 18 grams of water and then the number of molecules in one drop.

(1) Number of drops in 1 gram of water ____________________drops

(2) Number of drops in 1 mole (18 g) of water ____________________drops

(3) Number of molecules in 1 mole (18 g) of water ________________molecules

(4) Number of molecules in 1 drop of water ________________molecules

## EXPERIMENT 3 -- TEACHER'S GUIDE

## SCIENTIFIC NOTATION AND INDIRECT MEASUREMENT

### GENERAL TOPIC: III. SCIENTIFIC NOTATION AND SIGNIFICANT FIGURES

Introduction: Go over significant figures and scientific notation in detail with the students. Help them to learn the meteric relationships in making calculations. For instance, they should be able to readily do the following conversions.

$$1 \text{ meter} = 100 \text{ cm} = 0.001 \text{ km}$$

$$775 \text{ mL} = 0.775\text{L} = 775000 \ \mu\text{L}$$

$$0.625 \text{ mg} = 625 \ \mu\text{g} = 6.25 \times 10^{-4} \text{ g} = 6.25 \times 10^{-7} \text{ kg}$$

$$84.2 \text{ cm} = 842 \text{ mm} = 0.842 \text{ m}$$

If necessary have them work as follows:

$$0.625 \text{ mg} = \frac{0.625 \text{ mg} \times 10^6 \ \mu\text{g/g}}{10^3 \text{ mg/g}} = 625 \ \mu\text{g}$$

$$84.2 \text{ cm} = 84.2 \text{ cm} \times \frac{10^3 \text{ mm/m}}{10^2 \text{ cm/m}} = 842 \text{ mm}$$

The students should also understand how to read a ruler to the proper number of significant figures based on the units of the ruler. It is also important that the students understand how to use the balance and report the number of significant figures from its use.

Note: The ream of paper may be taken for weighing to a location that has a suitable balance or scales. Suggestions: hospitals, drug stores or grocery stores.

Sample Data and Calculations: The data given below for sections "C", "D", "F" is for illustration in doing the necessary calculations. Your students' data will vary according to the type of paper used in section "C" and the medicine droppers used for section "D".

### A. Significant Figures

Express the following in the proper number of significant figures.

(1) 25.66 + 38.205 = 63.86

(2) 46.22 - 2.2183 = 44.00

(3) 25.66 x 38.235 = 981.2

(4) 46.2 ÷ 2.7843 = 16.6

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Scientific Notation and Indirect Measurement, page 2
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## B. Scientific Notation

Express the following in scientific notation keeping the correct number of significant figures.

(1) 0.00234 = $2.34 \times 10^{-3}$

(2) 52300 = $5.23 \times 10^{4}$

(3) 0.000200 = $2.00 \times 10^{-4}$

(4) $3.62 \times 10^{-3} - 1.75 \times 10^{-2}$ = $-1.39 \times 10^{2}$

(5) $0.004304000 + 2.18 \times 10^{-3}$ = $6.48 \times 10^{-3}$

(6) $2.50 \times 10^{-7} \times 3.11 \times 10^{2}$ = $7.78 \times 10^{-5}$

(7) $(8.1 \times 10^{16}) \div (2.77 \times 10^{4})$ = $2.3 \times 10^{21}$

## C. Determining Thickness and Mass of a Single Sheet of Paper

(Data will vary according to the type of paper used.)

| | | |
|---|---|---|
| (1) | Thickness of 500 sheets | 4.92 cm |
| (2) | Thickness of 1 sheet in cm<br>(4.92 cm ÷ 500 sheets) | 0.00984 cm |
| (3) | Thickness of 1 sheet in µm | 98.4 µm |
| (4) | Mass of 500 sheets | 918.3 g |
| (5) | Mass of 1 sheet<br>(918.3 g ÷ 500 sheets) | 1.837 g |
| (6) | Mass of 1 sheet in mg | 1837 mg |
| (7) | Length of paper | 27.94 cm |
| (8) | Width of paper | 21.59 cm |
| (9) | Area of 1 sheet of paper<br>(29.94 cm x 21.59 cm) | 603.2 $\text{cm}^2$ |
| (10) | Mass of 1 $\text{cm}^2$ of 1 sheet of paper in g (1.837 g ÷ 603.2 $\text{cm}^2$) | $3.045 \times 10^{-3}$ $\text{g/cm}^2$ |
| (11) | Mass of 1 $\text{cm}^2$ of 1 sheet of paper in mg | 3.045 mg |
| (12) | Mass of 1 $\text{cm}^3$ of paper in g | 0.309 $\text{g/cm}^3$ |

$[918.3\ \text{g} \div (603.2\ \text{cm}^2 \times 4.92\ \text{cm})]$

or

$[1.837\ \text{g} \div (603.2\ \text{cm}^2 \times 9.84 \times 10^{-3}\ \text{cm})]$

or

$[(3.045 \times 10^{-3})\ \text{g} \div (1\ \text{cm}^2 \times 9.84 \times 10^{-3}\ \text{cm})]$

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Scientific Notation and Indirect Measurement, page 3
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### D. Determining the Volume of One Drop of Water

(Data will vary with the type of medicine dropper.)

(1) Number of drops in 3.0 grams of water. 60. drops

(2) Volume of 3.0 grams of water 3.0 mL
(Assume the density of water is 1.0 g/mL)

(3) Volume of 1 drop of water in mL 0.050 mL
(3.0 mL ÷ 60. drops)

(4) Volume of 1 drop of water in μL 50. μL

(5) Mass of 1 drop of water in mg 50. mg
(3.00 g ÷ 60 drops = 0.050 g = 50. mg)

(6) Mass of 1 drop of water in μg $5.0 \times 10^4$ μg

### E. Determing the Number of Molecules in One Drop of Water

(1) Number of drops in 1 gram of water 20. drops
(60. drops ÷ 3.0 g = 20. drops/g)

(2) Number of drops in 1 mole (18 g) of water
(20 drops/g x 18 g/mole = 360 drops) 360 drops

(3) Number of molecules in 1 mole of water
$6.023 \times 10^{23}$ molecules

(4) Number of molecules in 1 drop of water
[$6.023 \times 10^{23}$ molecules/mole ÷ (360 drop/mole)]
$1.7 \times 10^{21}$ molecules

Acknowledgement: This experiment is based in part on an experiment by McGill, Bradbury, and Sigler in Chemistry Guide and Laboratory Activities, Lyons and Carnahan, 1962.

## EXPERIMENT 4

## SYMBOLS AND FORMULAS

### GENERAL TOPIC: IV. NAMING, SYMBOLS AND FORMULAS

**Introduction:** All chemical compounds must be electrically neutral. Ionic compounds are electrically neutral even though they are made up of a combination of positively and negatively charged ions. You can learn the formulas of several ionic compounds and also write the names of the compounds by following the procedure given below.

**Objectives:**

1. To illustrate that chemical compounds must be electrically neutral.
2. To show the ratio in which ions combine to form a chemical compound.

**Materials and Equipment:**

cardboard or construction paper

graph paper

scissors

**Chemicals:** None

**Procedure:** On page 3 of this experiment you will find diagrams rep-resenting various ions. Make several copies of each ion diagram by tracing or copying on graph paper. Place these copies on cardboard or heavy paper, and cut out with scissors (CAUTION...HANDLE SCISSORS WITH CARE).

Arrange the positive ions on the left and the negative ions on the right. Refer to the list given below each diagram for symbols or formulas and names applicable to each of your cutouts.

Begin putting together diagrams of positive and negative ions so that all notches are filled by projections and no projections are left out, thus forming rectangles.

To write the formulas of the compounds you have made, count the number of positive ions of one kind that you have used and put down the symbol or formula of that ion followed by a subscript number showing the number of ions of that kind you have used. Immediately after the subscript write the symbol or formula of any other positive ion with its subscript. When you have written symbols or formulas for all the kinds of positive ions in your compound, write the symbols or formulas of the negative ions in the same way, along with their subscripts.

Experiment 4
Symbols and Formulas, page 2

Parentheses are used around the formula of any ion that consists of two or more different kinds of atoms if more than one of that ion is present.

Data and Results: After you have written the formula of the compound, write the name of the compound.

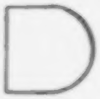
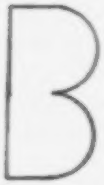
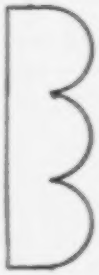
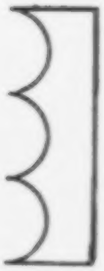
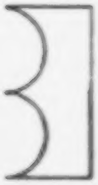
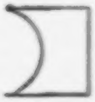
Questions:

1. How many negative ions with a charge of one are needed to combine with a positive ion with a charge of one?
2. How many positive ions with a charge of one are needed to combine with a negative ion with a charge of two?
3. How many positive ions with a charge of one are needed to combine with a negative ion with a charge of three?
4. How many negative ions with a charge of one are needed to combine with a positive ion with a charge of two?
5. How many negative ions with a charge of two, and positive ions with a charge of three, are needed to combine?

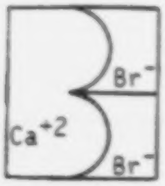
Further Study:

1. Are there additional compounds that could be constructed and named using the above technique?
2. Write formulas for some additional compounds.

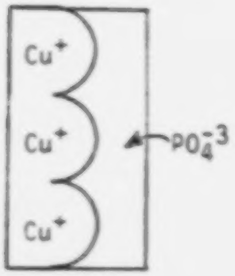
Experiment 4
Symbols and Formulas, page 3

| +1 | +2 | +3 | -3 | -2 | -1 |
|---|---|---|---|---|---|
|  |  |  |  |  |  |
| $Li^+$<br>lithium | | | | | $NO_3^-$<br>nitrate |
| $K^+$<br>potassium | | | | | $C_2H_3O_2^-$<br>acetate |
| $NH_4^+$<br>ammonium | $Ca^{+2}$<br>calcium | | | $O^{-2}$<br>oxide | $OH^-$<br>hydroxide |
| $Cu^+$<br>copper (I) | $Fe^{+2}$<br>iron (II) | | | $SO_3^{-2}$<br>sulfite | $Br^-$<br>bromide |
| | $Cu^{+2}$<br>copper (II) | $Al^{+3}$<br>aluminum | $N^{-3}$<br>nitride | $CO_3^{-2}$<br>carbonate | |
| | $Hg_2^{+2}$<br>mercury (I) | $Cr^{+3}$<br>chromium | $PO_4^{-3}$<br>phosphate | | |

EXAMPLES



$CaBr_2$
calcium bromide



$Cu_3PO_4$
copper(I) phosphate

# EXPERIMENT 4 -- TEACHER'S GUIDE

## SYMBOLS AND FORMULAS

### GENERAL TOPIC: IV. NAMING, SYMBOLS AND FORMULAS

**Objectives:**

1. Students will be able to show that all chemical compounds must be electrically neutral.
2. Students will be able to show how ions combine to form compounds.
3. Students will be able to identify several symbols for elements and symbols or formulas for ions.

**Pre-laboratory Discussion:**

Ionic compounds are composed of a combination of positively and negatively charged ions. These ions combine in a ratio that makes the resulting compound electrically neutral. This ratio is indicated in the formula of a compound by using appropriate subscripts when multiples of one or both ions are required.

**Material and Equipment:**

cardboard or construction paper

graph paper

scissors

**Notes on Teaching:**

1. It is assumed that the subjects of writing formulas and naming compounds have been discussed. In this exercise, after students have completed writing the formula for a compound, they are to write its name.

2. If some of the students finish before others, they could be given a handout on naming compounds when given the formulas and writing formulas when given names of compounds. This handout could be based upon formulas and names of compounds mentioned in the textbook.

**Answers to Questions:**

1. 1
2. 2
3. 3
4. 2
5. 3 negative ions and 2 positive ions.

## EXPERIMENT 5

## CHEMICAL EQUATIONS

## GENERAL TOPIC: V. EQUATIONS

Introduction: A chemical equation is chemical "shorthand" representing a chemical change. To write an equation correctly all reactants and products must be known, all formulas and symbols must be written correctly, and there must be the same number of each kind of atom on both sides of the equation.

Most chemical reactions fall into four general classes; in a synthesis reaction, two or more substances combine to form a more complex substance. A complex substance breaks down to form two or more simpler substances in a decomposition reaction. One substance is displaced by another in a single displacement reaction. A double displacement reaction is one in which the positive and negative ions of two compounds are interchanged. The general forms and examples of each of these reaction types are as follows:

Synthesis:

$$A + X \dashrightarrow AX$$

$$2\ Fe(s) + O_2(g) \dashrightarrow 2\ FeO(s)$$

Decomposition:

$$AX \dashrightarrow A + X$$

$$CaCO_3(s) \dashrightarrow CaO(s) + CO_2(g)$$

Single Displacement:

$$A + BX \dashrightarrow AX + B$$

$$Mg(s) + 2\ HCl(aq) \dashrightarrow MgCl_2(aq) + H_2(g)$$

$$Y + BX \dashrightarrow BY + X$$

$$Cl_2(g) + 2\ NaBr(aq) \dashrightarrow 2\ NaCl(aq) + Br_2(g)$$

Double Displacement:

$$AX + BY \dashrightarrow AY + BX$$

$$AgNO_3(aq) + HCl(aq) \dashrightarrow AgCl(s) + HNO_3(aq)$$

Experiment 5
Chemical Equations, Page 2

Objective: To perform four types of chemical reactions, to observe the formation of the products of these reactions, and to write a balanced equation for each reaction.

Materials and Equipment:

Bunsen burner
clamp
100 mL graduated cylinder
10 mL graduated cylinder
a file
wire gauze
litmus paper, red (for teacher's demonstration)
test tube holder or clamp
ring
ring stand
tongs
large test tube (20 X 200 mm)
test tube (16 X 150 mm)
watch glass (for teacher's demonstration)

Chemicals:

0.1 M aqueous solution of barium chloride, $BaCl_2$
6 M aqueous solution of hydrochloric acid, HCl
magnesium, ribbon (for teacher's demonstration)
0.1 M aqueous solution of magnesium sulfate, $MgSO_4$ (Epsom salts)
two 1984 (or later) pennies
sugar, $C_{12}H_{22}O_{11}$ (use granulated or sugar cubes)

Procedure: Prepare a table as shown below for recording data and observe both the reactants and the products.

| Reactants | Data/Observations | Products |
|---|---|---|
| | | |

A. Synthesis: DEMONSTRATION BY TEACHER

1. The teacher will obtain 5 cm of magnesium (Mg) ribbon.

2. Observe the ribbon.

3. The teacher will grasp the ribbon with tongs and hold in the burner flame. CAUTION: DO NOT LOOK DIRECTLY AT THE BURNING MAGNESIUM. The product of this reaction will be collected on a watch glass. Observe the product. The teacher will add a few drops of water and test with red litmus paper. Observe the litmus paper to see if it changes color. (The soupy residue may be wiped up with a paper towel and discarded in the trash can.)

B. Decomposition:

1. Obtain 1 to 2 grams of sugar, $C_{11}H_{22}O_{11}$, and place in a 15 X 150 mm test tube.

2. Support the test tube by attaching a clamp at the open end of the test tube. Angle the test tube at a $30^{o}$ angle from the horizontal.

3. Heat only the bottom of the test tube using a low flame.

4. As the sugar melts and eventually bubbles, observe the condensation of liquid on the inner wall near the opening of the test tube.

5. When all the liquid has been driven off, continue to heat until the sugar has changed color completely from white to black indicating the formation of carbon.

6. Allow to cool and observe the residue.

7. Discard this test tube in the waste can.

C. Single Displacement:

1. Obtain two 1984 (or later) pennies. Measure and record the mass of each penny.

2. Hold a flat or triangular file firmly, then grip a penny and stroke it on the file, cutting several notches around the rim exposing a tiny bit of the zinc filler. Turn the coin over and expose the zinc opposite each notch. Repeat this procedure with the second penny.

Experiment 5
Chemical Equations, Page 4

3. Obtain 40 mL of 6 M HCl in a 100 mL graduated cylinder. CAUTION: HCl CAUSES BURNS; AVOID SKIN CONTACT.

4 Allow coins to slide down the side of the cylinder into the acid. CAUTION: KEEP AWAY FROM OPEN FLAME since a flammable gas is being formed. Observe the reaction. Place the cylinder under the hood or where there is good ventilation to carry the hydrogen outdoors. At the end of this class period observe the reaction and let stand until class the next day.

5. At the beginning at the next day's class observe the pennies in the acid.

6. Carefully pour the acid off into a beaker in the sink. Then flush the contents of the beaker down the drain with a large quantity of running water. CAUTION: HCl CAUSES BURNS: AVOID SKIN CONTACT. Rinse the pennies thoroughly and dry. Weigh each penny and record the mass of each.

D. Double Displacement:

1. Pour 5 mL of 0.1 M barium chloride solution in a large test tube.

2. Add 5 mL of 0.1 M magnesium sulfate solution and observe.

Questions:

1. What substance reacted with the magnesium ribbon?

2. Often metal oxides will turn red litmus blue when moist. Write the chemical formula for the product formed when magnesium burned.

3. What was the liquid which formed on the funnel when the sugar was heated?

4. What was the flammable gas formed when the zinc in the penny reacted with HCl?

5. Determine the % copper in a 1984 (or later) penny.

6. Using the activity series, explain why the copper in the penny did not react with the acid although zinc did react.

7. Which of the two products of the double displacement reaction ($BaCl_2$ + $MgSO_4$) is soluble?

8. Write the chemical formula for the white precipitate formed in the double displacement reaction.

9. Write a balanced equation for each reaction.

## EXPERIMENT 5 -- TEACHER'S GUIDE

## CHEMICAL EQUATIONS

## GENERAL TOPIC: V. EQUATIONS

Objectives: Upon completion of this experiment when students are asked to demonstrate, diagram, or respond orally or on a written test, they will be able to:

1. identify four general types of chemical reactions and give an example of each.
2. predict the products of these types of reactions.
3. write a balanced equation for each type of reaction.

Preparation of Solutions:

| Reagent | Concentration | Amount Needed |
|---|---|---|
| 6 M HCl | 513 mL/L* | 40 mL/student |
| 0.1 M $BaSO_4$ | 23.3 g/L | 5 mL/student |
| 0.1 M $MgSO_4$ | 12.0 g/L | 5 mL/student |

*513 mL of concentrated HCl per liter of solution

Pre-laboratory Discussion:

1. There are several changes in appearance which signal that a chemical change has taken place: (1) color change, (2) formation of a precipitate, (3) formation of a gas, and (4) temperature change -- either a liberation or absorption of energy. (A reaction which releases energy is exothermic; one which absorbs energy is endothermic.)
2. Make careful observations and record all data and observations.
3. CAUTION: AVOID CONTACT WITH HCl, A STRONG ACID. STUDENTS SHOULD NOTIFY TEACHER IMMEDIATELY IF ACID IS SPILLED. IF SPILLED ON THE SKIN, WASH THE AFFECTED AREA WITH PLENTY OF WATER AND USE A PASTE MADE OF SODIUM HYDROGEN CARBONATE, $NaHCO_3$ (BAKING SODA) AND WATER TO NEUTRALIZE ANY RESIDUAL ACID.

4. Place the graduated cylinder containing the penny and hydrochloric acid under the vent hood and keep it away from open flames. Hydrochloric acid vapors are corrosive and toxic. Do not leave the pennies in HCl(aq) more than 24 hours.

5. Allow the glassware to cool completely before touching.

6. Do not place any solid materials in sink.

7. Flush acid down drain with plenty of running water.

Data and Calculations:

| | |
|---|---|
| Original mass of 1984 penny | 2.6 g |
| Mass of penny after reaction (mass of copper) | 0.3 g |
| Mass of zinc (original mass - mass of copper) | 2.3 g |

$$\%\text{ copper in penny} = \frac{\text{mass of copper}}{\text{original mass of penny}} \times 100$$

$$= \frac{0.3\text{ g}}{2.6\text{ g}} \times 100 = 11.5\%$$

Answers to Questions:

1. $O_2(g)$
2. $MgO(s)$
3. $H_2O(l)$
4. $H_2(g)$
5. $\%\text{ Cu in 1984 penny} = \frac{0.3\text{ g}}{2.6\text{ g}} \times 100 = 11.5\%$

6. The activity series has been determined experimentally to give the activity of metals and of hydrogen with metals. Activity decreases moving down the series. An element will replace a less active element (one below it on the series) but will not replace a more active element (one above it on the series).

ACTIVITY SERIES

potassium
calcium
sodium
magnesium
aluminum
zinc
iron
nickel
tin
lead
hydrogen
copper
mercury
silver
platinum
gold

Zinc is more active than hydrogen (zinc is above hydrogen on the series) and will replace hydrogen in hydrochloric acid. Copper is less active than hydrogen (copper is below hydrogen on the series) and will not replace hydrogen in hydrochloric acid. Therefore, the zinc is removed from the penny and the copper is left.

7. $MgCl_2$
8. $BaSO_4$

9. Balanced chemical equation for each reaction:

Synthesis:

$$2\ Mg(s) + O_2(g) \text{ ---> } 2\ MgO(s)$$

Decomposition:

$$C_{12}H_{22}O_{11}(s) \text{ ---> } 12\ C(s) + 11\ H_2O(g)$$

Experiment 5
Chemical Equations, page 4
Teacher's Guide

Single Displacement:

$$Zn(s) + 2\ HCl(aq) \text{ ---> } ZnCl_2(aq) + H_2(g)$$

Double Displacement:

$$BaCl_2(aq) + MgSO_4(aq) \text{ ---> } BaSO_4(s) + MgCl_2(aq)$$

Waste Disposal: Directions for waste disposal are given throughout the experimental procedure except for the chemicals involved in the double decomposition reaction. Of the four chemicals involved -- $BaCl_2(aq)$, $MgSO_4(aq)$, $BaSO_4(s)$, and $MgCl_2(aq)$ -- only $BaCl_2(aq)$ is dangerous. This solution is quite poisonous. Add a 20% excess of $MgSO_4(aq)$ to the products of the reaction to make certain that all barium ions have been precipitated as insoluble and, therefore, harmless $BaSO_4(s)$. Let the mixture stand overnight, then decant off the liquid, flushing it down the drain with running water. Empty the soupy solid on waste paper, wrap it up, and discard it in the trash can.

# EXPERIMENT 6

## EXPERIMENTAL STOICHIOMETRY

## GENERAL TOPIC: VI. STOICHIOMETRY AND THE MOLE CONCEPT

**Introduction:** The concept of stoichiometry is usually introduced using balanced equations and the mole concept. Stoichiometry is the quantitative relationship between reactants and products in a balanced chemical equation. This experiment is designed to show the student that an equation has an experimental basis. The chemistry involves the reactions of acids with bases.

**Objective:** To show by experiment that the quantities of materials used in chemical reactions are related to balanced chemical equations.

**Material and Equipment:**

one 50 mL beaker

one small graduated cylinder (5 mL, or 10 mL)

one medicine dropper

six 20 x 150 mm test tubes

**Chemicals:**

1.0 M aqueous solution of hydrochloric acid, HCl

phenolphthalein indicator

3.0 M aqueous solution of sodium hydroxide, NaOH

1.0 M aqueous solution of sulfuric acid, $H_2SO_4$

distilled water (for use in preparation of solutions)

CAUTION: BE CAREFUL WITH THESE REAGENTS; THEY ARE CORROSIVE.

**Procedure:**

1. Into a small graduated cylinder, measure 2.0 mL of 1.0 M hydrochloric acid and pour into a test tube. Add two drops of phenolphthalein indicator.

2. Obtain about 10 mL of 3.0 M sodium hydroxide in your small beaker.

3. Using the medicine dropper, add the sodium hydroxide solution (base), one drop at a time, to the acid in the test tube. Be sure to hold the dropper vertically to ensure that the drop size is uniform. Shake the test tube from side to side after each drop of base is added. (DO NOT SHAKE UP AND DOWN; THIS MAY CAUSE SPLASHING.) Add base until a color change occurs.

Experiment 6
Experimental Stoichiometry, page 2

At this point the acid is neutralized by the base. Record the total number of drops of base required under Data and Results.

4. Repeat the procedure (steps 1 to 3) using a new 2.0 mL sample of hydrochloric acid. There should not be a difference of more than two drops in the two runs. Be sure to use the same dropper in all runs.

5. Determine the average number of drops of base required to neutralize 2.0 mL of 1.0 M hydrochloric acid.

6. As in steps (1) to (5), find the average number of drops of base required to neutralize:
   (a) 4.0 mL of 1.0 M acid.
   (b) 2.0 mL of 1.0 M sulfuric acid.

Data and Results: Make a table for the number of drops of base used as follows:

| | Acid sample | Run 1 | Run 2 | Average |
|---|---|---|---|---|
| #1 | 2.0 mL of 1.0 M HCl | | | |
| #2 | 4.0 mL of 1.0 M HCl | | | |
| #3 | 2.0 mL of 1.0 M $H_2SO_4$ | | | |

Questions:

1. When the volume of hydrochloric acid was doubled, what change was observed in the number of drops of sodium hydroxide solution required to completely react with the acid?

2. What was the relationship between the number of drops of base required to react with equal volumes of the two different acids (hydrochloric acid and sulfuric acid)?

3. Write balanced chemical equations for the reactions that you have studied.

4. (a) Try to relate the answer for question 2 (above) to the equations in question 3.

   (b) Can you see why one acid required more base to neutralize it than the other acid did?

   (c) Explain.

# EXPERIMENT 6 -- TEACHER'S GUIDE

## EXPERIMENTAL STOICHIOMETRY

## GENERAL TOPIC: VI. STOICHIOMETRY AND THE MOLE CONCEPT

Objectives:

1. To show that an acid or base is neutralized when the amount of $H^+$ ($H_3O^+$) ions from the acid equals the amount of $OH^-$ ions from the base.

2. To show that for a monoprotic acid (such as HCl or $HNO_3$), one mole of acid involves one mole of $H^+$ ($H_3O^+$) ions and that for a diprotic acid (such as $H_2SO_4$), one mole of acid involves two moles of $H^+$ ($H_3O^+$) ions.

3. To show that when the concentration of the acid is increased, the amount of base consumed is also increased accordingly.

4. To illustrate that since $H_2SO_4$ is a diprotic acid, a given concentration of $H_2SO_4$ will consume twice as much NaOH as an equal molar concentration of HCl.

5. To illustrate that a balanced equation for an acid-base reaction should indicate the mole ratio between the acid and the base.

6. To illustrate the utility of setting up a table of values when making quantitative measurements.

7. To show that the products of a reaction of an acid with a base are a salt and water.

Pre-laboratory Discussion: The concept of stoichiometry is usually introduced using balanced equations and the mole concept. Many students have trouble with this approach. This experiment is designed to show the student that an equation has an experimental basis.

Balance any of these equations that are not balanced:

$Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + H_2O$ (or HOH)

$Ba(OH)_2 + HCl \rightarrow BaCl_2 + H_2O$ (or HOH)

$NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$ (or HOH)

What type of reaction does each equation represent?

Experiment 6
Experimental Stoichiometry, page 2
Teacher's Guide

Preparation of Solutions: One liter of each solution should be sufficient for 50 students.

1M HCl: Measure carefully 86 mL of concentrated hydrochloric acid and add this to 200 mL of distilled water contained in a 1 L volumetric flask. Swirl the flask to stir, and allow the solution to cool to room temperature. When cool, dilute to the mark with distilled water. Stopper and allow to stand for a few minutes; then invert a few times to ensure complete mixing. Transfer to a storage bottle and label.

1M $H_2SO_4$: Carefully measure 56 mL of concentrated sulfuric acid and add this to 400 mL of distilled water contained in an 800 mL beaker. CAUTION: DILUTING CONCENTRATED SULFURIC ACID RELEASES A LARGE AMOUNT OF HEAT. ADD THE ACID SLOWLY WITH CONSTANT STIRRING WITH A GLASS ROD. REMEMBER, ALWAYS ADD ACID TO WATER. MAKE SURE YOU ALLOW THIS SOLUTION TO REACH ROOM TEMPERATURE BEFORE PROCEEDING. Add the solution to a 1 L volumetric flask. Add 100 mL of distilled water to the beaker and stir. Add this wash liquid to the volumetric flask. Swirl to stir. Dilute to the mark with distilled water. Stopper and allow to stand for a few minutes; then invert several times to insure complete mixing. Transfer to a storage container and label.

3M NaOH: SOLID SODIUM HYDROXIDE IS VERY HYGROSCOPIC (RAPIDLY ABSORBS MOISTURE FROM THE AIR) AND CORROSIVE. ANY PELLETS DROPPED SHOULD BE CLEANED UP IMMEDIATELY, BUT DO NOT TOUCH THEM WITH THE BARE HAND. Into a 600 mL beaker weigh 120 g. of solid reagent grade sodium hydroxide. Add 400 mL of distilled water and stir the mixture. CAUTION: DISSOLVING SOLID SODIUM HYDROXIDE IN WATER RELEASES A LARGE AMOUNT OF HEAT. THE VAPOR ABOVE THE SOLUTION IS CORROSIVE TO MEMBRANES. BE VERY CAREFUL. MAKE SURE YOU ALLOW THIS SOLUTION TO COOL TO ROOM TEMPERATURE BEFORE PROCEEDING. When cool, transfer to a 1 L volumetric flask. Add 100 mL of distilled water to the beaker and stir; then add this wash liquid to the volumetric flask. Swirl to mix. Dilute to the mark with distilled water. Stopper and invert several times to insure complete mixing. Transfer to a PLASTIC STORAGE CONTAINER and label.

Phenolphthalein Solution: Dissolve 1 g of phenolphthalein in 50 mL of ethyl alcohol and add 50 mL of water.

Experiment 6
Experimental Stoichiometry, page 3
Teacher's Guide

**Notes on teaching:**

1. It is essential that the concentrations of the two acids (HCl and $H_2SO_4$) be equal but not necessarily 1.0 M.

2. The concentration of HCl decreases with time. A new solution should be prepared each year.

3. The NaOH should be stored only in plastic bottles; but may be placed in glass if not kept too long.

4. The concentration of the NaOH can be approximately three times greater than the concentration of the acids without affecting the number of drops required.

5. Two different students may have different data because of variations in droppers. For a given student the data should be consistent as long as the same dropper is used.

6. This experiment introduces a simplified form of titration. It allows students to consider some of the underlying principles without being faced with difficult techniques and calculations.

7. Your students need to understand solution concentrations, be familiar with simple reaction types, and be able to balance equations by inspection.

**Sample Data:**

| ACID SAMPLE | RUN 1 | RUN 2 | AVERAGE |
|---|---|---|---|
| 2.0 mL HCl | 10 | 11 | 10.5 |
| 4.0 mL HCl | 21 | 19 | 20 |
| 2.0 mL $H_2SO_4$ | 21 | 20 | 20.5 |

**Answers to Questions:**

1. It required twice as many drops for neutralization of 4.0 mL of 1 M hydrochloric acid as it did for 2.0 mL of 1 M hydrochloric acid.

2. It required twice as many drops for the neutralization of 2.0 mL of 1 M sulfuric acid as it did 2.0 mL of 1 M hydrochloric acid.

3. (a) $HCl + NaOH \rightarrow NaCl + HOH$

   (b) $H_2SO_4 + 2\ NaOH \rightarrow Na_2SO_4 + 2\ HOH$

Experiment 6
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4. Since we used equal volumes of the same concentration of acids, one might think that we should have the same number of acid molecules present. However the sulfuric acid required twice as many drops of base for neutralization. The explanation is that a sulfuric acid molecule is different from a hydrochloric acid molecule in that the sulfuric acid molecule furnishes two $H^+$ ions; whereas a hydrochloric acid molecule furnishes one $H^+$ ion. Therefore, one molecule of sulfuric acid can react with two molecules of sodium hydroxide; whereas one hydrochloric acid molecule reacts with only one molecule of sodium hydroxide.

   In looking at the balanced equations one sees that this information is denoted by the relative numbers of molecules on the reactant side.

Waste Disposal: Each solution can be neutralized to a phenolphthalein endpoint and flushed down the drain with water.

Acknowledgement: This experiment is based on one published by Michael J. Webb in J. Chem. Ed., 58, 192 (1981).

Answers to Pre-laboratory Discussion Questions:

$$Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2\ H_2O$$

$$Ba(OH)_2 + 2\ HCl \rightarrow BaCl_2 + 2\ H_2O$$

$$2\ NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2\ H_2O$$

All of these reactions are acid-base reactions yielding a salt and water.

## EXPERIMENT 7

## INVESTIGATION OF HARD AND SOFT WATER

### GENERAL TOPIC: VII. DESCRIPTIVE CHEMISTRY

Introduction: The mineral concentrations in water have an important influence on the uses of that water. Water which is high in calcium, and/or magnesium and/or iron content is called "hard water". Hard water is a problem in many areas and causes problems in the home because of its tendency to form insoluble precipitates with soap. In this experiment, we will study some of the properties of hard water.

Objective: To define "hard water" and to learn some of the problems caused by hard water.

Material and Equipment:

balance, 0.01 g sensitivity
3 test tubes (15 mm X 125 mm)
test tube rack

Chemicals:

5% aqueous solution of calcium chloride, $CaCl_2$
calcium sulfate, $CaSO_4$ (solid)
liquid detergent
5% aqueous solution of magnesium chloride, $MgCl_2$
solid soap
sodium carbonate, $Na_2CO_3$ (solid)
distilled water
hard water

Procedure: There are no special safety precautions required for this experiment.

1. Arrange three test tubes in a test-tube rack.
2. Half fill the first test tube with distilled water, the second with tap water, and the third with hard water.
3. Add pea-sized samples of solid soap to each tube, cork the test tubes, and shake them well.

Record your observations for each tube.

4. Half fill three test tubes with tap water. Add a small chip of soap to each.

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Investigation of Hard and Soft Water, page 2

5. To the first tube add 5 drops of 5% calcium chloride solution. To the second tube add 5 drops of 5% magnesium chloride solution. Do not add anything to the third test tube.

6. Cork and shake each tube.

Record your observations.

7. Half fill two test tubes with tap water. Add 0.5 g. of solid calcium sulfate to each tube.

8. To the first tube add a few drops of detergent; then cork and shake the tube. To the second tube add 0.5 g. of solid sodium carbonate. Cork and shake the tube. Then add a few drops of detergent to the tube containing sodium carbonate. Shake the tube again, record and explain your observations.

Data and Results:

Distilled water and solid soap: ____________________________

____________________________________________

____________________________________________

Tap water and solid soap: ____________________________

____________________________________________

____________________________________________

Hard water and solid soap: ____________________________

____________________________________________

____________________________________________

Calcium chloride solution and soap: ____________________________

____________________________________________

____________________________________________

Magnesium chloride solution and soap: ____________________________

____________________________________________

____________________________________________

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Investigation of Hard and Soft Water, page 3

Tap water and solid soap: ______________________________

______________________________________________________

______________________________________________________

Calcium sulfate and liquid detergent: ___________________

______________________________________________________

______________________________________________________

Calcium sulfate, sodium carbonate and liquid detergent: ___

______________________________________________________

______________________________________________________

______________________________________________________

Questions:

1. What ions are usually associated with "hard water"?

2. In this experiment, what was added to hard water which eliminated these ions from solution?

3. What effect was shown by the liquid detergent?

# EXPERIMENT 7 -- TEACHER'S GUIDE

## INVESTIGATION OF HARD AND SOFT WATER

## GENERAL TOPIC: VII. DESCRIPTIVE CHEMISTRY

**Objectives:** The student should be able to:

1. identify ions that cause "hard water."
2. show with experimental observations why "hard water" is undesirable.
3. identify substances that can be added to "hard water" to make it soft.

**Pre-laboratory Discussion:** As water moves along or beneath the surface of the earth, it dissolves minerals as well as matter from decaying plants and animals. Recall that minerals (salts) are ionic in nature, and that ions have either positive or negative charges. The principal positive ions in natural water are sodium, potassium, calcium, magnesium, and sometimes iron. The negative ions are usually sulfate, hydrogen carbonate (bicarbonate), and chloride. The presence of calcium, magnesium, and iron salts makes the use of soap inconvenient. Water containing calcium, magnesium and/or iron ions is called "hard water". These positive ions react with negative ions of soap to form insoluble organic salts (with the appearance of a scum) which cling to clothes, leaving them dingy in appearance. This scum is also observed as the familiar "bathtub ring".

There are two major ways to avoid the problem of hard water. A softening agent like sodium carbonate can be added. The carbonate ion reacts with the positive ions to form an insoluble carbonate salt (an inorganic salt), thereby preventing the formation of the undesirable scum which is an organic salt.

The second way to avoid this problem is to use a synthetic detergent. Molecules of a synthetic detergent are enough like those of soap to have the same cleansing action. However, the negative ions of a detergent do not form insoluble organic salts with the positive ions found in hard water; therefore, no scum is formed.

**Preparation of Solutions:**

| | |
|---|---|
| 5% $CaCl_2$ Solution: | Add 5 grams $CaCl_2 \cdot 2H_2O$ to 95 mL of distilled water. |
| 5% $MgCl_2$ Solution: | Add 5 grams of $MgCl_2 \cdot 6H_2O$ to 95 mL of distilled water. |

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Hard Water Solution: Add 2 grams of $CaCl_2 \cdot 2H_2O$ and 2 grams of $MgCl_2 \cdot 6H_2O$ to 96 mL of distilled water.

Sample Data and Results:

Distilled water and solid soap: Soap dissolves with good sudsing action.

Tap water and solid soap: Unless water is very hard, the reaction will probably be the same as the distilled water.

Hard water and solid soap: Undissolved precipitate forms and good suds are not formed.

Calcium chloride solution and soap: Undissolved precipitate forms and good suds are not formed.

Magnesium chloride solution and soap: Undissolved precipitate forms and good suds are not formed.

Tap water and solid soap: Unless the water is very hard, it will probably react the same as distilled water.

Calcium sulfate and liquid detergent: The calcium ions may prevent very good suds formation, but no undissolved precipitate forms.

Calcium sulfate, sodium carbonate and liquid detergent: Carbonate ions combine with the calcium ions to form insoluble calcium carbonate. Removal of calcium ions from solution allows good suds formation.

Answers to Questions: See experiment for questions.

1. calcium and magnesium ions
2. sodium carbonate
3. No insoluble precipitate is formed.

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Post-laboratory Discussion:

1. Why is hard water undesirable?

   Answer: Because calcium and/or magnesium and/or iron ions present in hard water react with the negative ions of soap to form insoluble organic salts which appear as an undesirable scum.

2. Is our water hard or soft?

   Answer: Varies, depending on the locality.

3. Does anyone in class use water softening techniques in the home? It would be very helpful to demonstrate the use of a commercial water softening apparatus. Perhaps a local distributor of these products would arrange a demonstration for you.

4. How do modern detergents allow us to clean our clothes better than the pioneers could clean theirs with their homemade soap?

   Answer: Detergents are similar to soap in their cleaning ability, but the negative ions of a detergent do not form insoluble organic salts with the positive ions found in hard water. Consequently, no scum is formed.

Waste Disposal: There are no significant waste disposal problems with this experiment. All solutions can be flushed down the sink with plenty of water. Solids may be wrapped in paper and placed in the trash can.

# EXPERIMENT 8

# IONS AND THEIR REACTIONS

## GENERAL TOPIC: VII. DESCRIPTIVE CHEMISTRY

Introduction: A large number of chemical compounds when dissolved in water will exist as ions in solution. Consider the dissolution of potassium nitrate, $KNO_3$.

$$KNO_3(\text{solid}) \dashrightarrow K^+(aq) + NO_3^-(aq)$$

Potassium nitrate is an ionic solid and when dissolved in water will come apart to form hydrated ions as shown above. The parenthetical (aq) means that the ion is in solution and has molecules of water attached to it. A mixing of the aqueous solutions of two compounds such as,

$Pb(NO_3)_2$ and KI,

would thus be mixing the hydrated ions resulting from the dissolving of the two compounds.

You will be given 9 solutions. These solutions consist of the following compounds dissolved in water:

| Compound | Chemical Formula | Ions In Solution |
|---|---|---|
| barium chloride | $BaCl_2$ | $Ba^{+2}$, $Cl^-$ |
| copper(II) sulfate | $CuSO_4$ | $Cu^{+2}$, $SO_4^{-2}$ |
| iron(III) chloride | $FeCl_3$ | $Fe^{+3}$, $Cl^-$ |
| hydrochloric acid | $HCl(aq)$ | $H^+$, $Cl^-$ |
| lead(II) nitrate | $Pb(NO_3)_2$ | $Pb^{+2}$, $NO_3^-$ |
| silver nitrate | $AgNO_3$ | $Ag^+$, $NO_3^-$ |
| sodium carbonate | $Na_2CO_3$ | $Na^+$, $CO_3^{-2}$ |
| sodium hydroxide | $NaOH$ | $Na^+$, $OH^-$ |
| sodium phosphate | $Na_3PO_4$ | $Na^+$, $PO_4^{-3}$ |

Objectives:

1. To demonstrate the importance of making accurate observations.
2. To organize and analyze a large amount of data.
3. To correlate observations with chemical reactions.

4. To determine the chemical composition of some unknown solutions.

5. To study the behavior of ions in solution.

Material and Equipment:

9 bottles (glass or polyethylene)

9 medicine droppers or dropping pipets

A spot plate

Chemicals:

1 M solutions of:

$BaCl_2$

$CuSO_4$

$FeCl_3$

$HCl$

$Pb(NO_3)_2$

$AgNO_3$

$Na_2CO_3$

$NaOH$

$Na_3PO_4$

distilled or deionized water (for use in preparation of solutions)

Procedure:

(SAFETY NOTE: WEAR EYE PROTECTION AT ALL TIMES DURING THIS EXPERIMENT. TREAT ALL CHEMICALS AS CAUSTIC AND TOXIC. IF THE SOLUTIONS GET ON YOU, WASH THEM OFF WITH PLENTY OF WATER. DO NOT INGEST ANY OF THESE SOLUTIONS OR EAT AFTER HANDLING THESE CHEMICALS WITHOUT WASHING YOUR HANDS. DO NOT PIPET USING MOUTH SUCTION.)

1. Obtain a set of "known" solutions and a spot test plate.

2. Place 3 drops of one solution in each of the eight depressions on the spot test plate.

3. Add, dropwise, 3 drops each of the other reagents.

4. Record complete observations such as color changes, precipitate formation, gas evolution, etc.

5. Rinse the plate and repeat with the next solution on the spot test plate in the same fashion.

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6. When you have completed this for all nine "known" solutions, show your recorded observations to your instructor and an unknown will be assigned to you.

7. Repeat the same sequence of tests on the "unknown" solution as you did for the "known" solutions.

8. Check your observations for the "unknown" solution against the observations for the "known" solutions.

9. Identify your unknown solution.

Results: Record your observations on the following chart according to the nine test solutions. Note that there is no reaction when a solution is mixed with itself; hence,"X" is written in these blocks.

Chart 1.

| | #1 | #2 | #3 | #4 | #5 | #6 | #7 | #8 | #9 |
|---|---|---|---|---|---|---|---|---|---|
| #1 | X | | | | | | | | |
| #2 | | X | | | | | | | |
| #3 | | | X | | | | | | |
| #4 | | | | X | | | | | |
| #5 | | | | | X | | | | |
| #6 | | | | | | X | | | |
| #7 | | | | | | | X | | |
| #8 | | | | | | | | X | |
| #9 | | | | | | | | | X |

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Chart 2.

| IONS | $Ba^{+2}$ | $Cu^{+2}$ | $Fe^{+3}$ | $H^+$ | $Pb^{+2}$ | $Ag^+$ | $Na^+$ |
|---|---|---|---|---|---|---|---|
| $SO_4^{-2}$ | | | | | | | |
| $NO_3^-$ | | | | | | | |
| $CO_3^{-2}$ | | | | | | | |
| $OH^-$ | | | | $H_2O$ | | | |
| $PO_4^{-3}$ | | | | | | | |

Questions: 1. Using the charts 1 & 2, write in the blanks below the identity of the ions and the compound used to prepare the unknown solutions.

| Unknown # | Cation | Anion | Compound in the solution |
|---|---|---|---|
| #1 | ______ | _____ | __________________________ |
| #2 | ______ | _____ | __________________________ |
| #3 | ______ | _____ | __________________________ |

2. Use the solubility rules (given on the next page) to predict the reactions, if any, which will occur when solutions containing the cations listed horizontally below are mixed with solutions containing the anions listed vertically below. Write "NR" if an insoluble compound is not formed and "PPT" if an insoluble compound is formed. All spaces should have either an "NR" or an "PPT" when they have been filled in.

## SOLUBILITY RULES

I. Acids

All acids are soluble.

II. Bases

All hydroxides are insoluble except those of Group IA ions, Group IIA heavy ions, $Sr^{+2}$ and $Ba^{+2}$, and the ammonium, $NH_4^+$, ion.

III. Salts

1. All salts containing $NH_4^+$ or Group IA ions or the ammonium, $NH_4^+$, ion are soluble.
2. All nitrate, $NO_3^-$, salts are soluble.
3. All chloride, $Cl^-$, salts are soluble except those of silver, $Ag^+$; lead(II), $Pb^{+2}$; and mercury(I), $Hg_2^{+2}$.
4. All sulfate, $SO_4^{-2}$, salts are soluble except those of $Ag^+$, $Pb^{+2}$, $Ca^{+2}$, $Sr^{+2}$, and $Ba^{+2}$.
5. All carbonate, $CO_3^{-2}$, salts are insoluble except those of Group IA cations and ammonium, $NH_4^+$, ion.
6. All phosphate, $PO_4^{-3}$, salts are insoluble except those of Group IA ions and the ammonium, $NH_4^+$, ion.
7. All sulfide, $S^{-2}$, salts are insoluble except those of Group IA ions, Group IIA ions, and the ammonium, $NH_4^+$, ion.

## EXPERIMENT 8 -- TEACHER'S GUIDE

## IONS AND THEIR REACTIONS

### GENERAL TOPIC: VII. DESCRIPTIVE CHEMISTRY

Objectives: The student should be able to:

1. describe what is meant by the term "precipitation reaction".
2. make careful and accurate observations.
3. correlate observations with chemical reactions.
4. organize and analyze a large amount of data.
5. apply logical reasoning (deduction) in order to identify the chemical composition of some unknown solutions.

Pre-laboratory Discussion: Precipitation reactions take place when positive ions (cations) and negative ions (anions) from soluble compounds form a solid compound in aqueous solution. Ions in solid ionic compounds are held together by coulombic attraction of unlike charges. The solubility rules tell you which solids will dissolve.

Point out to the students that you cannot make 1.0 molar solutions of AgCl, $Fe(OH)_3$ or other insoluble salts.

Make sure that the students understand that a solution cannot contain only cations or only anions. The coulombic forces of repulsion between ions of the same charge require that oppositely charged ions must be present to result in an electrically neutral solution. Thus both types of ions must be present so that the total of the positive charges equals the total of the negative charges. In order to make an aqueous solution containing desired cations or anions, one must dissolve in water a soluble ionic compound containing the desired ions.

The mixing of two different ionic solutions results in two possibilities:

1. A clear solution is formed (no precipitate or cloudy mixture). This happens when all ionic combinations are soluble in water.
2. A precipitate or cloudy solution is formed.

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The solubility rules and the exercise sheet can be assigned either prior to this experiment or after the laboratory observations. The data chart organizes the observations in a clear and useful array. Students should be familiar with the known solutions before they are given the unknown solution(s). The response form for Question 1 is designed to assist the student in identifying the component ions in the unknown(s). This exercise involves deductive reasoning and is both interesting and challenging to the student.

**Preparation of Chemicals:** Use only distilled or deionized water in preparing the solutions for this experiment. Store the silver nitrate, $AgNO_3(aq)$ solution in brown glass bottles. Prepare 0.10 molar solutions of the reagents in 1 liter, 500 mL, or 250 mL quantities (depending on the amount needed). The chart below should be useful. (Elsewhere in this handbook you will find detailed instructions for preparing solutions of a given molarity.) BE CAREFUL WHEN PREPARING THESE SOLUTIONS! SOME ARE VERY CORROSIVE. SOME ARE QUITE POISONOUS.

| Solution # | Reagent chemical | Volume | Moles | Grams |
|---|---|---|---|---|
| 1. | $BaCl_2$ | 1 liter | 0.10 | 20.82 |
| | | 500 mL | 0.05 | 10.41 |
| | | 250 mL | 0.025 | 5.20 |
| 2. | $CuSO_4 \cdot 5H_2O$ | 1 liter | 0.10 | 24.46 |
| | | 500 mL | 0.05 | 12.23 |
| | | 250 mL | 0.025 | 6.12 |
| 3. | $FeCl_3$ | 1 liter | 0.10 | 16.22 |
| | | 500 mL | 0.05 | 8.11 |
| | | 250 mL | 0.025 | 4.06 |
| 4. | Conc. HCl<br>(12 M) | 1 liter | 0.10 | 8.33 mL |
| | | 500 mL | 0.05 | 4.17 mL |
| | | 250 mL | 0.025 | 2.08 mL |
| 5. | $Pb(NO_3)_2$ | 1 liter | 0.10 | 33.12 |
| | | 500 mL | 0.05 | 16.56 |
| | | 250 mL | 0.025 | 8.28 |
| 6. | $AgNO_3$ | 1 liter | 0.10 | 16.99 |
| | | 500 mL | 0.05 | 8.50 |
| | | 250 mL | 0.025 | 4.25 |

| Solution # | Reagent chemical | Volume | Moles | Grams |
|---|---|---|---|---|
| 7. | $Na_2CO_3$ | 1 liter<br>500 mL<br>250 mL | 0.10<br>0.05<br>0.025 | 10.60<br>5.30<br>2.65 |
| 8. | NaOH | 1 liter<br>500 mL<br>250 mL | 0.10<br>0.05<br>0.025 | 4.00<br>2.00<br>1.00 |
| 9. | $Na_3PO_4$ | 1 liter<br>500 mL<br>250 mL | 0.10<br>0.05<br>0.025 | 16.39<br>8.20<br>4.10 |

Samples to be distributed as "unknowns" are selected from these solutions and are marked as unknown samples.

Suggestions for dispensing these solutions are given below.

**Notes on Techniques and Hazards:**

SAFETY NOTE: INSIST THAT EYE PROTECTION BE WORN AT ALL TIMES IN THE LABORATORY. THESE REAGENT SOLUTIONS SHOULD BE TREATED AS THE SEVERE POISONS THAT MOST OF THEM ARE. BE SURE TO HAVE THE STUDENTS WASH THEIR HANDS CAREFULLY AFTER THIS LABORATORY PRIOR TO EATING. SILVER NITRATE IS NOT ONLY VERY CORROSIVE BUT IS DECOMPOSED BY LIGHT AND PERMANENTLY STAINS CLOTHES, SHOES AND SKIN. THE BLACK STAIN ON SKIN WILL WEAR AWAY WITHIN SEVERAL WEEKS.

A dilute solution of sodium thiosulfate, $Na_2S_2O_3$, can be used to remove silver nitrate stains before the silver ions are reduced to silver metal (the black stain) by the light in the room. Often the black silver stains will appear after the laboratory is long over.

There are several ways of dispensing the reagents for this experiment. It is important to avoid cross contamination of solutions or the results will be worthless. The best way to supply the reagents is through the use of 50 mL or 100 mL plastic dropping bottles. A set of nine separate dropping bottles obviates the need for medicine droppers and will avoid cross contamination of the reagents. These bottles may be placed in a rack or a small box.

An alternative method is to use one liter or 500 mL stock bottles and pipettes in each. The student fills test tubes and then uses medicine droppers to dispense the reagents from the test tubes. Unless the test tubes are well marked it is easy to make mistakes as to which solution is being tested.

It is very likely that cross contamination will occur at an unsupervised stocking station, especially if the students lay the pipettes down on the bench top after transferring the solution to the test tube. For this reason take the caps off the stock bottles and place a pipette equipped with a rubber bulb in each stock solution bottle. DO NOT ALLOW THE STUDENTS TO PIPETTE USING MOUTH SUCTION. Always use a rubber bulb.

Answers to Questions:

1. The results in the chart below are listed as PPT (precipitate), GAS (bubbles evolved), X (same reagent - no reaction), NR (no reaction). The precipitates are white in color except for $Fe(OH)_3$ (red) and $Ag_2O$ (muddy brown) in basic solutions.

CHART 1

| | #1 | #2 | #3 | #4 | #5 | #6 | #7 | #8 | #9 |
|---|---|---|---|---|---|---|---|---|---|
| #1 | X | PPT | NR | NR | PPT | PPT | PPT | NR | PPT |
| #2 | PPT | X | NR | NR | PPT | PPT | PPT | PPT | PPT |
| #3 | NR | NR | X | NR | PPT | PPT | GAS | PPT | PPT |
| #4 | NR | NR | NR | X | PPT | PPT | GAS | NR | NR |
| #5 | PPT | NR | PPT | PPT | X | NR | PPT | PPT | PPT |
| #6 | PPT | PPT | PPT | PPT | NR | X | PPT | PPT | PPT |
| #7 | PPT | PPT | PPT/GAS | GAS | PPT | PPT | X | NR | NR |
| #8 | NR | PPT | PPT | NR | PPT | PPT | NR | X | NR |
| #9 | PPT | PPT | PPT | NR | PPT | PPT | NR | NR | X |

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2. CHART 2

| | $Ba^{+2}$ | $Cu^{+2}$ | $Fe^{+3}$ | $H^+$ | $Pb^{+2}$ | $Ag^+$ | $Na^+$ |
|---|---|---|---|---|---|---|---|
| $SO_4^{-2}$ | PPT | NR | NR | NR | PPT | PPT | NR |
| $NO_3^-$ | NR | NR | NR | NR | NR | NR | NR |
| $CO_3^{+2}$ | PPT | PPT | PPT | NR | PPT | PPT | NR |
| $OH^-$ | NR | PPT | PPT | $H_2O$ | PPT | PPT | NR |
| $PO_4^{-3}$ | PPT | PPT | PPT | NR | PPT | PPT | NR |
| $Cl^-$ | NR | NR | NR | NR | PPT | PPT | NR |

Additional Post-laboratory Exercises:

1. Indicate whether or not the following combinations will form a precipitate. If a precipitate forms, then write the balanced equation for the reaction which occurs. If there is no reaction, then write "NO REACTION".

   (a) $BaCl_2(aq) + Na_2CO_3(aq)$ --> ?

   (b) $CuSO_4(aq) + NaOH(aq)$ --> ?

2. $K_3PO_4$ is (soluble) (insoluble). Circle the correct answer.

3. $PbCl_2$ is (soluble) (insoluble). Circle the correct answer.

4. Iron(III) chloride and ammonium nitrate solutions react upon being mixed. (true) (false). Circle one response.

5. $Ba_3(PO_4)_2$ dissolves in water. (true) (false). Circle one response.

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Answers:

1. (a) $BaCl_2(aq) + Na_2CO_3(aq) \dashrightarrow BaCO_3(s) + 2\ NaCl(aq)$

   (b) $CuSO_4(aq) + NaOH(aq) \dashrightarrow Cu(OH)_2(s) + Na_2SO_4(Aq)$

2. soluble

3. insoluble

4. false

5. false

Waste Disposal: Since use of the spot test plates requires so little of the reagents, there is no disposal problem. The spot plates can be rinsed with water and wiped clean with paper towels between tests. The bottled reagents may not be disposed of as easily. Most solutions have a long shelf life and may be saved to use again. The NaOH solution can be poured down the drain with a large quantity of running water. $AgNO_3$ does not store well in solution and in time decomposes due to light. Barium, lead, and silver compounds and solutions are extremely toxic. Don't prepare excess quantities of the solutions. When you must dispose of these toxic solutions, use the procedure described under "General Method for Disposing of Solutions of Toxic Inorganic Salts" found near the end of Chapter 5.

## EXPERIMENT 9

## ENERGY OF A CHEMICAL CHANGE

### GENERAL TOPIC: VIII. ENERGY AND CHEMICAL CHANGE

**Introduction:** In this experiment an acid and a base will be reacted in water solution. The heat liberated by the reaction increases the temperature of the water solution. This increase in temperature and the specific heat of water will be used to calculate the heat liberated by the reaction.

1. Heat gained by the = (Specific heat) X (grams of) X (temperature)
   water solution (water) (solution) (rise)

   or Heat = (specific heat) x (mass) x $\Delta t$
   (where $\Delta t$ = final temperature - initial temperature)

2. heat gained by the water solution = heat liberated by the reaction

**Objective:** To experimentally measure the energy involved in a chemical change.

**Material and Equipment:**
balance, 0.01 g sensitivity
12 oz. styrofoam cups
100 mL graduated cylinder
250 mL Erlenmeyer flask
thermometer, 1°C graduation

**Chemicals:**
0.6 M aqueous solution of sodium hydroxide, NaOH
0.3 M aqueous solution of sulfuric acid, $H_2SO_4$

**Precautions:** BOTH SODIUM HYDROXIDE SOLUTION AND SULFURIC ACID ARE STRONGLY CORROSIVE. DO NOT USE THESE CHEMICALS WITHOUT ADEQUATE EYE PROTECTION. SHOULD ETIHER SOLUTION GET ON YOUR SKIN, WASH IT OFF IMMEDIATELY WITH A LARGE QUANTITY OF RUNNING WATER.

**Procedure:**

1. Construct a solution calorimeter by using two styrofoam cups, one inside the other.
2. Carefully measure 50 mL of 0.6 M NaOH using a graduated cylinder and pour into the inner styrofoam cup.
3. Measure the temperature of this solution and record.
4. Measure 50 mL of 0.3 M $H_2SO_4$ and add to the inner styrofoam cup with stirring. Record the temperature of this

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Energy of a Chemical Change, page 2

solution every 25 seconds until the solution reaches a maximum temperature. Record this maximum temperature as the final temperature.

Data and Results:

mL of 0.6 M NaOH used ______________ mL

Initial temperature of solution ______________ °C

mL of 0.3 M $H_2SO_4$ used ______________ mL

Total volume of solution ______________ mL

Final temperature of solution ______________ °C

Heat gained by solution ______________ J

Heat of reaction per mole of water produced ______________ kJ

Questions:

1. Using the data from this experiment calculate the heat (in Joules) released by this reaction.

2. Write a balanced chemical equation for the reaction studied.

3. Knowing the amounts of the solutions used, calculate the moles of water produced by the reaction.

4. Calculate the heat of reaction per mole of water produced.

# EXPERIMENT 9 -- TEACHER'S GUIDE

## ENERGY OF A CHEMICAL CHANGE

## GENERAL TOPIC: VIII. ENERGY AND CHEMICAL CHANGE

Objectives:

1. The student should understand the use of a styrofoam cup calorimeter.
2. The student should be able to relate the terms exothermic and endothermic to energy changes in reactions.
3. The student should be able to calculate the energy change as measured in a styrofoam cup calorimeter.
4. The student should be able to calculate the energy change per mole of reactant or product.

Pre-laboratory Discussion:

1. Make sure your students are familiar with the terms exothermic and endothermic.
2. Go over the use of the equation:

   Heat = (specific heat) X ( mass) X ($\Delta$t)
3. Review stoichoimetric calculations.
4. Make sure your students are familiar with the $\Delta H$ notation for heat of reaction: The sign of $\Delta H$ is negative for an exothermic reaction; positive for an endothermic reaction.

Preparation of Solutions:

NaOH: See the Teacher's Guide in Experimental Stoichiometry for detailed instructions. To prepare 1 L of 0.3 M NaOH you will use 12 g of NaOH.

$H_2SO_4$: See the Teacher's Guide in Experimental Stoichiometry for detailed instructions. To prepare 1 L of 0.6 M $H_2SO_4$ you will use 33 mL of concentrated $H_2SO_4$.

Notes on Teaching:

1. In this experiment you can assume the total mass of solution to be equal to the total volume. This is because the concentrations are so small the densities are essentially 1.0 g/mL -- the density of water.
2. Make sure that the solutions are at room temperature before using.

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Energy of a Chemical Change, page 2
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3. The specific heat of water is 4.18 J/g°C.

4. The students should read the thermometers as accurately as possible.

**Sample Data and Calculations:**

50 mL 0.6 M NaOH used

50 mL 0.3 M $H_2SO_4$ used

initial temperature = 23.0°C

final temperature = 27.0°C

heat = 4.184 J/g°C X 100 g solution X (27.0 - 23.0)°C

heat = 1670 Joules

We are reacting 0.03 mole of NaOH with 0.015 mole $H_2SO_4$. Thus, we make 0.03 mole of water. The balanced equation is

$$2\ NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2\ H_2O$$

Thus: $$\frac{\text{Heat generated}}{2\ \text{mole}\ H_2O} = \frac{1670\ \text{joules}}{0.03\ \text{mole}\ H_2O}$$

heat generated = 111 kJ

Since the reaction is exothermic,
heat of reaction = -111 kJ

**Answers to Questions:**

1. See the Sample Data and Calculations for illustrative example.
2. $2\ NaOH(aq) + H_2SO_4 \longrightarrow Na_2SO_4(aq) + 2\ H_2O(l)$
3. 0.03 mole of water
4. $\Delta H = -111$ kJ

**Post-laboratory Discussion:**

1. Make sure the student understands the convention for $\Delta H$ (heat of reaction) notation.

2. The correct value for the heat of reaction is -111 kJ. Choose some student results and calculate a percent error. If you have an error of less than 15%, the experiment is satisfactory.

3. You could relate the use of a calorimeter for energy measurements to the energy value of foods. This would be a natural lead-in to a discussion of nutrition, weight loss, and exercise.

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Waste Disposal: Used solutions may be poured down the drain. Unused solutions may be mixed for neutralization and then poured down the drain and flushed with ample water.

## EXPERIMENT 10

## THE MOLECULAR MASS OF A GAS

### GENERAL TOPIC: IX. GASES AND GAS LAWS

**Introduction:** Using Avogadro's Hypothesis we know that the molar volume of a gas, at 273 K and 1 atm., is 22.4 liters. We can use this information to experimentally determine the molecular mass of a gas. In this experiment the gas used is butane, commonly found in pocket lighters.

**Precaution:** THE LIGHTER IS USED ONLY AS A SOURCE OF BUTANE. IT IS NOT TO BE IGNITED UNDER ANY CIRCUMSTANCES.

**Objective:** To determine the molecular mass of butane gas.

**Material and Equipment:**

balance, 0.01 g sensitivity

butane lighter

100 mL graduated cylinder

250 mL Erlenmeyer flask

marking pencil, (or tape)

glass plate (or a watch glass)

thermometer, 1°C graduation

trough

**Procedure:**

1. Fill your 250 mL Erlenmeyer flask completely with water. Place a small glass plate on top of the flask.

2. Fill the trough or a container designated by your teacher about 3/4 full with water. Keeping the glass plate in place on the Erlenmeyer, carefully invert the flask and place it in the water contained in the trough. No air bubbles should enter the flask. Remove the glass plate.

3. Obtain a butane lighter and wipe it dry and clean with a paper towel.

4. Determine the mass of the lighter to ± 0.01 g and record this mass in the data table.

5. Carefully hold the lighter under the water. Be sure that the gas opening is just below the mouth of the flask. Press the release lever to allow gas from the lighter to enter the flask.

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The Molecular Mass of a Gas, page 2

6. Collect about 200 mL of gas. Do not remove the gas-filled flask from the water.

7. Remove the lighter, dry it, and place it to one side.

8. You must now equalize the pressure inside the flask with the atmospheric pressure. To do this we raise or lower the flask until the water level inside is the same as the water level outside.

9. Mark this water level on the flask. Remove the flask and fill with water until this mark is reached. Using your graduated cylinder, you can now measure the volume of the flask occupied by the gas. Record this volume in the data table.

10. Reweigh the lighter to ± 0.01 g. Record this mass.

11. Measure the temperature of the water in the trough. Record. Also be sure to obtain and record the barometric pressure.

Data and Results:

Data Table - Butane Collection

Mass of lighter before use ___________g

Mass of lighter after use ___________g

Mass of butane ___________g

Volume of gas collected ___________mL

Temperature ___________°C

Pressure ___________mm Hg

Questions:

1. Using data collected in laboratory, calculate the molecular mass of butane gas.

2. Find the accepted value for the molecular mass of butane from a standard reference and calculate the percent error for your determination.

3. Could this method be used for any gas? What are some of the limitations?

# EXPERIMENT 10 -- TEACHER'S GUIDE

## THE MOLECULAR MASS OF A GAS

## GENERAL TOPIC: IX. GASES AND GAS LAWS

Objectives: The student should be able to:

1. collect a gas over water.
2. make calculations using the combined gas laws.
3. apply the concept of molar volume.
4. calculate the molecular mass of butane using experimental data.

Pre-laboratory Discussion: Make sure that your students understand the combined gas law and the concept of molar volume.

Notes on Teaching:

1. The difference in the mass of the lighter before and after release of the gas will be small. This then requires that a large volume of gas be collected for reasonable accuracy. It is very important that accurate mass determinations be made. SHOW BY CALCULATION HOW A SMALL ERROR IN MEASUREMENT, 1 mL or 0.1 g, WILL AFFECT THE ANSWER.
2. Any disposable butane lighter will work. To minimize the possibility of ignition you should remove the flint and striker. Plug the hole left by the flint with modeling clay (or chewing gum).
3. Demonstrate how to equalize the pressure by raising or lowering the water levels.
4. A plastic container from your local discount store can be substituted for the trough. Some people have had great success in using plastic baby bottles for collection vessels.
5. If you don't have a barometer, call your local weather service and ask for the barometric pressure. It will probably be given in inches of Hg. Consult a handbook for the correct conversion factor.
6. Instead of a marking pencil you can use a piece of tape to mark the volume of gas collected.

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7. If your students are very careful, their results can be within 3-5% of the accepted value.

Sample Data:

Data Table - Butane Collection

| | |
|---|---|
| Mass of lighter before use | 21.34 g |
| Mass of lighter after use | 20.88 g |
| Mass of butane | 0.46 g |
| Volume of gas collected | 200 mL |
| Temperature (°C) | 19 °C |
| Pressure (mm Hg) | 756 mm Hg |

Answer to Questions:

1. (a) To determine the pressure of the butane gas, you must subtract the pressure due to the water vapor. Consult a table of vapor pressures for water at various temperatures (use your text book), and select the pressure that corresponds to the temperature that you measured. Subtract this vapor pressure of water from the barometric pressure.

$$P_{butane} = P_{barometric} - P_{water}$$

(b) Change the measured volume to standard conditions.

$$V_{STP} = V_{measured} \times \frac{P_{butane}}{760 \text{ mm}} \times \frac{273 \text{ K}}{(273 \text{ K} + \text{temp.}_{measured})}$$

(c) Set up a ratio.

$$\frac{\text{Mass of gas (g)}}{\text{Volume of gas at STP (mL)}} = \frac{\text{"X" of gas}}{\text{22,400 mL of gas}}$$

$$\text{"X"} = \frac{\text{(Mass of gas in grams)} \times \text{(22,400 mL gas)}}{\text{Volume of gas in milliliters}}$$

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The Molecular Mass of a Gas, page 3
Teacher's Guide

2. The accepted value is 58.1 g. To obtain percent error you take the difference in measured and accepted values divided by the accepted value, and the resulting fraction is multiplied by 100.

$$\% \text{ error} = \frac{\text{measured} - \text{accepted}}{\text{accepted}} \times 100$$

3. One of the limitations is solubility of the gas. A gas that is highly soluble in water could not be collected by displacement because it would dissolve in the water. Some gases of this type are hydrogen chloride and ammonia.

Waste Disposal: When all students have completed this experiment, the lighter should be taken outdoors, away from any source of ignition, and emptied of any remaining gas. Then the container can be discarded in a trash can.

Acknowledgement: This experiment is based on an experiment by Lee Summerlin in Chemistry of Common Substances, Silver Burdett, 1983.

## EXPERIMENT 11

## ISOTOPIC MASS VS. ATOMIC MASS

### GENERAL TOPIC: XI. ATOMIC THEORY

Introduction: If atoms were as large as beans they could be sorted, counted, and weighed. In this experiment the analogy is made between atoms and beans, with different types of beans representing different isotopes of the same element.

Objective: To show the relation between isotopic mass, isotopic abundance, and atomic mass.

Materials and Equipment:

balance, 0.01 g sensitivity

beakers (3) 400 mL

500 beans (three different types)

Procedure: You have before you a vessel containing the "atoms of an element". Can you see the different "isotopes" (beans)?

1. Sort the isotopes into groups. Record the total number of atoms (beans) in your sample as well as the number of each isotope (number of each kind of bean) on your data sheet.
2. On the data sheet, sketch a picture of each isotope. Try to show clearly the difference in each isotope.
3. Atomic and isotopic masses are actually relative masses and are based upon a reference standard, the carbon-12 isotope. In this experiment the absolute mass in grams of the isotopes will be determined. To determine the isotopic mass of each isotope we:
    (a) find the total mass of each isotope; record these masses.
    (b) divide the total mass of each isotope by the number of atoms in that particular sample to find the mass of a single atom of the isotope of interest, and hence its isotopic mass.
4. Determine the percent abundance of each isotope by dividing the number of atoms of the isotope by the total number of atoms in your sample. Repeat for each isotope. Record these data.

Experiment 11
Isotopic Mass vs Atomic Mass, page 2

5. Determine the atomic mass for your element based on the isotopic abundances and the isotopic masses.

ATOMIC MASS = % ISOTOPE #1 x MASS OF ONE ATOM OF ISOTOPE #1

+ % ISOTOPE #2 x MASS OF ONE ATOM OF ISOTOPE #2 +......n

Data and Results:

Total number of beans in container:______________

Isotope Information:

| TYPE OF BEAN | No. OF BEANS | MASS OF BEANS |
|---|---|---|
| (a) navy | ____________ | ____________ |
| (b) pinto | ____________ | ____________ |
| (c) blackeyed | ____________ | ____________ |

Questions:

1. How many atoms do you have of each isotope?

2. What is the isotopic mass of each isotope?

3. What is the percent abundance of each isotope?

4. What is the weighted atomic mass of the element?

5. Chlorine exists in nature as chlorine-35, isotopic mass 34.96885, percent abundance 75% and chlorine-37, isotopic mass 36.96590, percent abundance 25%. Show how to calculate the atomic mass of chlorine.

## EXPERIMENT 11 -- TEACHER'S GUIDE

## ISOTOPIC MASS VS. ATOMIC MASS

## GENERAL TOPIC: XI. ATOMIC THEORY

Pre-laboratory Assignment: It is suggested that this assignment be given the day before the experiment is done and then discuss the answers before beginning the experiment.

1. What is the basic atomic difference between isotopes of the same element?

2. If there are 150 navy beans, 32 pinto beans and 178 blackeyed peas in a container, what is the percent composition of the mixture by type of bean?

3. If your chemistry grade is broken down so that 70% of it is based on exams, 20% on lab reports, and 10% on homework, and your average scores (out of 100 points in each area) are:

| | |
|---|---|
| exams | 85 |
| labs | 75 |
| homework | 96 |

What will your weighted average score be?

Pre-laboratory Assignment Answers:

1. Different isotopes have different masses. This is due to the fact that the isotopes have a different number of neutrons in their nuclei.

2.

| TYPE OF BEAN | NUMBER | PERCENT COMPOSITON |
|---|---|---|
| (a) navy | 150 | $\frac{150}{360} \times 100 = 41.67\%$ |
| (b) pinto | 32 | $\frac{32}{360} \times 100 = 8.89\%$ |
| (c) blackeyed | $\frac{178}{360}$ | $\frac{178}{360} \times 100 = 49.44\%$ |

3. $0.70 \times 85 + 0.20 \times 75 + 0.10 \times 96 = 84.1$

Notes on Teaching:

1. You need about 500 beans per group. The beans can be of any type as long as you have three different types. Make

Experiment 11
Isotopic Mass vs Atomic Mass, page 2
Teacher's Guide

sure you obtain dried beans. A good way to store beans is to use plastic containers with tight fitting lids.

2. Stress that real isotopes would probably not have the visual differences that our beans have.

3. Also the student needs to understand that measuring the mass of a single atom in grams would be beyond the reach of our best experimentation.

Sample Data
and Results: Total number of beans in container : 636

Isotope Information:

| TYPE OF BEAN | No. OF BEANS | MASS OF BEANS |
|---|---|---|
| (a) navy | 128 | 42.758 |
| (b) pinto | 185 | 43.678 |
| (c) blackeyed | 323 | 58.757 |

Isotopic Masses $\frac{\text{(total mass of one kind of bean)}}{\text{(number of that kind of bean)}}$:

navy 42.758/128 = 0.33405

pinto 43.678/185 = 0.23610

blackeyed 58.757/323 = 0.18191

Isotopic Abundances $\frac{\text{(number of each type of bean)}}{\text{(total number of all beans)}}$:

navy 128/636 x 100 = 20.1%

pinto 185/636 x 100 = 29.1%

blackeyed 323/636 x 100 = 50.8%

Atomic Mass =
(0.201)x(0.33405)+(0.291)x(0.23610)+(0.508)x(0.18191) = 0.228

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Isotopic Mass vs Atomic Mass, page 3
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**Answers to Questions:**

1-4. Obviously the answers to questions 1-4 will vary with each student. The information presented above in "Sample Data and Results" should be an adequate guide for checking answers to questions 1-4.

5. Atomic Mass of Cl =

$$(0.75 \times 34.96885) + (0.25 \times 36.96590) = 35.468113$$

**Acknowledgement:**

The experiment is based on one published by G. C. Miller, *J. Chem. Ed.*, *58*, 507 (1981).

## EXPERIMENT 12

## ELECTRICAL CONDUCTIVITY AND CHEMICAL BONDING

### GENERAL TOPIC: XII. BONDING

Introduction: In this experiment, you will measure the electrical conductivity of several different pure liquids as well as some water solutions of compounds in order to predict the bond type in the substances tested. To understand electrical conductivity, we must first define electric current. An electric current is produced by the motion of positive and negative particles. The current is measured in terms of the total number of positive and negative charges passing a given point per second. In a solid, such as copper wire, only the negative particles (free electrons) are able to move. Electrical conductivity is a property which describes the ability of a substance to transmit an electric current.

There are two major types of compounds; ionic and covalent. The ionic type contains charged particles (ions); the covalent type does not. Electrical conductivity can be used as a test to distinguish between these types of compounds. If a solution of a compound conducts electricity, the dissolved compound is considered to be ionic. If the solution does not conduct, no ions are present, thus the compound is covalent. Materials that conduct electricity when melted or when dissolved in water are also called electrolytes.

In this experiment, if the pure liquid substance or solution to be tested conducts electricity, bubbles will be formed at the wires submerged in the liquid. Thus the appearance of bubbles is evidence of electrical conductivity.

Objective: To distinguish between ionic and covalent compounds by electrical conductivity.

Material and Equipment:

9-volt battery

conductivity cell

two paper clips

cork

4-dram vial

Chemicals:

calcium chloride, $CaCl_2$

lead acetate, $Pb(C_2H_3O_2)_2$

methyl or ethyl alcohol

potassium bromide, KBr

Experiment 12
Electrical Conductivity and Chemical Bonding, page 2

sodium chloride, NaCl

sucrose (sugar)

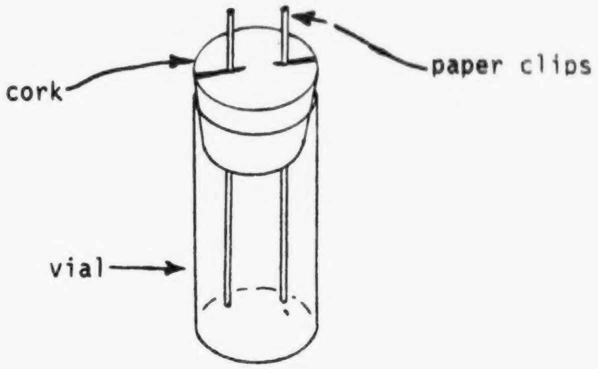
distilled or deionized water.

tap water

Procedures: Acquire the pure liquids and prepare the solutions listed in the chart on page 4 of this experiment. Using distilled or deionized water, prepare about 25 mL of solution of each compound: sodium chloride, sugar, calcium chloride, and potassium bromide. You can make each solution by adding about 0.5 g of the appropriate solid to 25 mL of distilled water with stirring. The alcohol test solutions are prepared in a similar fashion. When testing the alcohol, tap water, and distilled water they are used directly.

Set up a conductivity apparatus as illustrated below:

To test for conductivity



Prepare the conductivity cell by cutting two vertical slits on either side of a cork fitted to a 4 dram glass vial (or transparent plastic pill bottle from a pharmacy). Straighten two small paper clips and insert into the slits of the cork, as illustrated. Position the ends of the paper clips protruding from the top of the cork to match the electrode spacing of the 9-volt battery so that when the battery is touched to the apparatus, each terminals on the battery will touch a paper clip.

Fill the vial about 2/3 full of the test liquid and insert the cork and wire into the vial. Momentarily touch the battery electrodes to the leads of the cell and observe whether bubbles are generated about the wires in the solution. Less than a second is needed to observe the bubbles if formed.

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Electrical Conductivity and Chemical Bonding, page 3

(Caution) Use good laboratory techniques. Do not place unused portion of chemicals back into the reagent bottles. Check with the instructor concerning the proper disposal of all materials.

Data and Results: Use the table on page 4 of this experiment for recording your results. Rate the conductivity in each test material as good, poor, or none and record any other observations.

Questions:

1. What was the purpose of using both water and alcohol as solvents?

2. Explain any differences in your results for distilled and tap water.

3. On the basis of this experiment, what conclusions can you draw concerning the bonding in the four solids tested and two liquids tested?

Further Study:

1. Add 0.5 g (pinch) lead(II) acetate {$Pb(C_2H_3O_2)_2$} to 25 mL of water. Test the conductivity. Based on your observations, what can you predict about the bonding in this compound?

2. How would this compound fit into our bonding classification scheme?

Experiment 12
Electrical Conductivity and Chemical Bonding, page 4

| Test Material | Conductivity | Other Observations |
|---|---|---|
| Distilled Water | | |
| Tap Water | | |
| Alcohol | | |
| NaCl in water | | |
| NaCl in alcohol | | |
| Sucrose in water | | |
| Sucrose in alcohol | | |
| $CaCl_2$ in water | | |
| $CaCl_2$ in alcohol | | |
| KBr in water | | |
| KBr in alcohol | | |

## EXPERIMENT 12 -- TEACHER'S GUIDE

## ELECTRICAL CONDUCTIVITY AND CHEMICAL BONDING

### GENERAL TOPIC: XII. BONDING

**Objectives:** Students will be able to:

1. predict the bonding nature of each substance based on its conductivity.
2. identify a conductivity apparatus.
3. identify electrolytes and nonelectrolytes.

**Pre-laboratory Discussion:** Due to the nature of their chemical bond(s), some compounds when placed in water will simply "break apart" and produce ions in solution. This process is called dissociation. Compounds that can dissociate possess ionic bonds. Solutions of these compounds conduct electricity.

Other compounds when placed in water will react with the water to form ions. This process is called ionization and these compounds contain polar covalent bonds which break in aqueous solutions. Examples are HCl and HBr. Solutions of these compounds also will conduct electricity.

Because solutions of ionic compounds and solutions of polar covalent compounds will conduct electricity, both types of compounds are referred to as electrolytes.

There are other compounds that when placed in water will not produce ions; therefore, they will not conduct electricity. Compounds of this type are called nonelectrolytes. These are usually nonpolar covalently bonded compounds.

In this experiment, students are to observe the conductivity of several different liquid compounds and solutions in order to predict the bond type in the substances tested.

**Preparation of Solutions:** Solutions are to be prepared by adding 0.5 g of sodium chloride, sucrose, calcium sulfide, and potassium bromide to 15.0 mL of distilled water.

**Precautions:** Point out that the electrodes must not touch each other in the cell. Also point out that the 9-volt battery terminals should not be shorted out by allowing any metal to touch both terminals or else the batteries will be drained.

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Electrical Conductivity and Chemical Bonding, page 2
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**Hazards:** PURE ALCOHOLS ARE FLAMMABLE.

STUDENTS ARE NOT TO TASTE TEST ANY CHEMICALS IN THE LABORATORY AND SHOULD WASH THEIR HANDS THOROUGHLY AFTER FINISHING THE EXPERIMENT.

**Sample Data:** Rate the conductivity in each test material as good, poor, or NDC (no detectable conductivity)

| Test Material | Conductivity | Other Observation |
|---|---|---|
| solution of sucrose in water | NDC | clear solution |

**Answers to Questions:**

1. Water is a very polar solvent and ionic compounds dissociate into ions in water. Alcohol is not polar enough to cause dissociation of ionic compounds. Thus water solutions of ionic compounds conduct electricity but alcohol solutions will not.

2. Distilled water is not a conductor. Tap water has a small amount of salts dissolved in it and may conduct slightly. (Slow evolution of small bubbles on the electrodes may be seen.)

3. You should be able to distinguish clearly between the compounds tested and to know which possess ionic bonding. The covalent compounds do not conduct electricity.

**Waste Disposal:** The alcohol and all solutions with the exception of the lead acetate solution may be disposed of by pouring down the sink while flushing with large quantities of running water. Dispose of the solution of lead acetate by the procedure described under "General Method for Disposing of Solutions of Toxic Inorganic Salts" found near the end of Chapter 5.

## EXPERIMENT 13

## PREPARATION OF A STANDARD SOLUTION AND ITS USE IN COLORIMETRY

### GENERAL TOPIC: XIII. SOLUTIONS

**Introduction:** A standard solution (one of known concentration) of copper(II) sulfate will be prepared and then used to measure the concentration of an unknown copper(II) sulfate solution.

**Objective:** To show how a standard solution can be prepared and used.

**Materials and Equipment:**

- balance, 0.01 g sensitivity
- two small beakers
- plastic wash bottle
- graduated cylinder, preferably 25 or 50 mL
- medicine dropper
- sheet of white paper
- stirring rod
- two test tubes (15 mm x 125 mm)

**Chemicals:**

- copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O$ (solid)
- distilled or deionized water

**Procedure A:** Preparation of a standard solution:

1. Add a small amount of $CuSO_4 \cdot 5H_2O$ (between 1.0 - 1.5g) to a small dry beaker and weigh. Transfer the copper(II) sulfate to a second beaker being careful not to lose any.
2. Reweigh the first beaker. The difference between the mass of the beaker containing copper(II) sulfate and the mass of the empty beaker is the mass of the copper(II) sulfate.
3. Pour about 15 mL of distilled or deionized water into the beaker containing the copper(II) sulfate and stir with a stirring rod until all the solid has dissolved.
4. Pour the aqueous solution of copper(II) sulfate into a graduated cylinder. To minimize splashing, pour with the lip of the beaker touching a stirring rod held so that it extends into the mouth of the graduated cylinder.

Experiment 13
Preparation of a Standard Solution
and Its Use in Colorimetry, page 2

5. Using a plastic wash bottle, wash down the sides of the beaker that held the copper(II) sulfate solution with a small amount of water. Now add this washing to the graduated cylinder, again using the stirring rod to avoid splashing. Repeat the washing process once and transfer this rinse water also to the graduated cylinder.

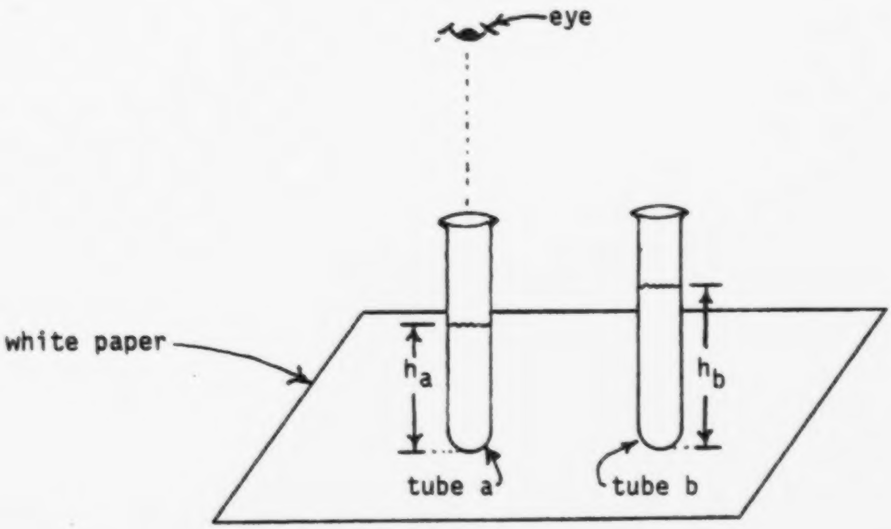
6. Now carefully add just enough water to the graduated cylinder to fill it to the 25 mL mark. Using the stirring rod, carefully stir the solution in the graduated cylinder to insure thorough mixing.

Procedure B: Measurement of the concentration of an unknown copper(II) sulfate solution by comparing its color intensity to that of the standard solution prepared earlier.

1. Fill a test tube (A) to a depth of about 7 cm. with an unknown copper(II) sulfate solution prepared by the instructor.

2. Add just enough of the standard copper(II) sulfate solution to a second test tube (B) to achieve the same color intensity in both of the two test tubes when viewed from above as they are held upright on a white sheet of paper.

3. Measure the heights of the solutions in the test tubes and record as $h_a$ and $h_b$.



Experiment 13
Preparation of a Standard Solution
and Its Use in Colorimetry, page 3

Calculations:

1. Calculate the molarity of your standard solution. (Moles of $CuSO_4 \cdot 5H_2O$ per liter)

2. Calculate the concentration of the unknown solution from the comparison of its color intensity to that of your standard. The solution with the greater concentration will have the lesser depth in its test tube when the color intensities appear to be the same.

   The comparison can be expressed mathematically as:

   $$\text{conc. A} \times h_a = \text{conc. B} \times h_b$$

EXPERIMENT 13
Preparation of a Standard Solution
and Its Use in Colorimetry, page 4

## Data Sheet:

Name(s)________________________________________

Procedure A:

Mass of beaker and $CuSO_4 \cdot 5H_2O$ ______________________

Mass of Beaker ______________________

Mass of $CuSO_4 \cdot 5H_2O$ ______________________

Volume of Solution (in liters) ______________________

Procedure B:

$h_a$ ______________________

$h_b$ ______________________

Calculated Data:

Molecular mass of $CuSO_4 \cdot 5H_2O$ ______________________

Moles of $CuSO_4 \cdot 5H_2O$ ______________________

Concentration of the standard solution (in moles per liter) ______________________

Concentration of the unknown solution (in moles per liter) ______________________

Questions:

1. What is the definition of a standard solution?

2. A 0.1 M solution of $CuSO_4 \cdot 5H_2O$ contains:

________moles of $CuSO_4 \cdot 5H_2O$ per liter.

________moles of $CuSO_4 \cdot 5H_2O$ per 100 mL.

________moles of $CuSO_4 \cdot 5H_2O$ per 25 mL.

## EXPERIMENT 13 -- TEACHER'S GUIDE

## PREPARATION OF A STANDARD SOLUTION AND ITS USE IN COLORIMETRY

### GENERAL TOPIC: XIII. SOLUTIONS

Objectives:

1. To prepare a copper(II) sulfate solution of known concentration.

2. Using the prepared solution as a standard, the concentrations of one or more unknown solutions will be determined by comparing the color intensities of the unknowns relative to that of the standard.

Preparation: Unknown Solution: Appropriate concentrations of unknown solutions can be prepared by dissolving an accurately weighed amount (25 to 50 g) of $CuSO_4 \cdot 5H_2O$ in enough water to make one liter of solution. The molar mass of $CuSO_4 \cdot 5H_2O$ is 250 g.

$$\text{Molarity} = \frac{\text{g } CuSO_4 \cdot 5H_2O}{\text{1 L solution}} \times \frac{\text{1 mole } CuSO_4 \cdot 5H_2O}{\text{250 g } CuSO_4 \cdot 5H_2O}$$

Notes on Techniques: The test tubes used must be of the same size. It would be preferable to use tubes with a flat bottom.

The higher the depth of standard and unknown solutions in the tubes, the less the error in the determination. This type of determination (known as a colorimetric determination) obviously can only be used with colored reagents such as copper sulfate. The test tubes are held upright side by side and the color comparison is made by looking straight down the top of the liquid against a white background. Good illumination is needed. The test tubes can be held with one hand and the unknown liquid is poured into the empty test tube beside the standard solution. A medicine dropper can be used to adjust the color intensity until a perfect match is obtained. Do not mix up the reagents while adding them to the test tubes.

Precaution: COPPER(II) SULFATE MUST NOT BE INGESTED SO THE STUDENTS SHOULD BE WARNED TO WASH THEIR HANDS AFTER COMPLETING THIS EXPERIMENT.

Experiment 13
Preparation of a Standard Solution
and Its Use in Colorimetry, page 2
Teacher's Guide

**Sample Data :**

Finding the concentration of a standard solution:

If 1.10 g of $CuSO_4 \cdot 5H_2O$ is dissolved in enough water to make 25 mL of solution,

Conc. of standard =

$$\frac{1.10\text{ g } CuSO_4 \cdot 5H_2O}{0.025\text{ L soln.}} \times \frac{1\text{ mole } CuSO_4 \cdot 5H_2O}{250\text{ g } CuSO_4 \cdot 5H_2O} = \frac{0.176\text{ moles } CuSO_4 \cdot 5H_2O}{\text{L Soln}}$$

$$= 0.176\text{ M}$$

Finding the concentration of an unknown solution:

A 7.4 cm depth of an unknown $CuSO_4$ solution appears to have the same color intensity as 3.2 cm of the standard.

conc. of standard x depth of standard =
conc. of unknown x depth of unknown

0.176 M x 3.2 cm = conc. of unknown x 7.4 cm

$$\text{conc. of unknown} = \frac{0.176\text{ M} \times 3.2\text{ cm}}{7.4\text{ cm}} = 0.076\text{ M}$$

**Answers to Questions:**

1. A solution of known concentration.

2. 0.1 moles of $CuSO_4 \cdot 5H_2O$ per liter

   0.01 moles of $CuSO_4 \cdot 5H_2O$ per 100 mL

   0.0025 moles of $CuSO_4 \cdot 5H_2O$ per 25 mL

**Waste Disposal:**

The solutions may be disposed of by flushing down the drain with a large quantity of running water.

## EXPERIMENT 14

## APPLYING "STRESS" TO A CHEMICAL SYSTEM IN EQUILIBRIUM

## GENERAL TOPIC: XIV. EQUILIBRIUM

Introduction: Stress will be applied to systems in equilibrium by (A) changing the concentration of one component, and (B) changing the temperature. The equilibrium shift will be observed.

Objective: To apply Le Chatelier's Principle in predicting observed equilibrium shifts

Materials and Equipment:

beaker, any size from 250 to 600 mL

graduated cylinder, 10 mL

medicine dropper

stirring rod

thermometer, 1°C graduation

2 test tubes (15 mm X 125 mm)

Chemicals:

concentrated hydrochloric acid, HCl

ice

potassium nitrate, $KNO_3$ (solid)

saturated aqueous solution of potassium nitrate, $KNO_3$

saturated aqueous solution of sodium chloride, NaCl

Procedure A:

1. Add 3 mL of saturated NaCl solution to a test tube.
2. Add 10 drops of concentrated HCl to the NaCl solution, stir, and record results. (NOTE: CONCENTRATED HCl SHOULD NOT COME INTO CONTACT WITH YOUR HANDS OR CLOTHING.)

Procedure B:

(a) To determine if dissolving solid $KNO_3$ in water is an endothermic or an exothermic process.

1. Fill a test tube half-full with water that is at room temperature.
2. Place a thermometer in the water and after one minute record the temperature to the nearest 1°C.
3. Leaving the thermometer in the water, add 0.5 to 1 g of solid $KNO_3$, stir slowly for one minute, and again record the temperature to the nearest 1°C.

Experiment 14
Applying "Stress" to a System in Equilbrium, Page 2

(b) The effect of temperature on solubility.

Add 5 mL of saturated $KNO_3$ solution to a test tube. Place the test tube in a beaker of ice water, stir for about one minute, and record the results.

**Data and Results:**

(A) Results of adding concentrated HCl to saturated NaCl solution.

(B) (a)

Temperature of water ________________________ °C

Temperature after $KNO_3$ dissolves______________ °C

(b)

Results of cooling the saturated solution of $KNO_3$.

**Questions:**

1. Referring to the equilibrium expression:

$Na^+Cl^-(s) \rightleftharpoons Na^+(aq) + Cl^-(aq)$

explain how Le Chatelier's Principle predicts the observed results in Procedure A.

Experiment 14
Applying "Stress" to a System in Equilibrium, page 3

2. Consider the results in part (a) of Procedure B. Is this solution process endothermic or exothermic? Explain your answer.

3. Consider part (b) of Procedure B. Write an equilibrium equation showing $KNO_3(s)$ going into and coming out of a saturated solution. Using your answer from question #2 (above), add the "heat" quantity to the correct side of the equation.

4. Explain how Le Chatelier's Principle predicts the observed results in part (b) of Proceure B.

## EXPERIMENT 14 -- TEACHER'S GUIDE
## APPLYING "STRESS" TO A SYSTEM IN EQUILIBRIUM
## GENERAL TOPIC: XIV. EQUILIBRIUM

Objectives: The student should:

1. observe how a change in concentration can shift the equilibrium process.
2. be able to determine experimentally if a solution process is endothermic or exothermic.
3. observe how a change of temperature can shift the equilibrium.
4. be able to use Le Chatelier's Principle in predicting shifts in equilibrium.

Pre-laboratory Discussion:

1. Make sure the students understand the dynamic nature of chemical equilibrium.
2. Explain Le Chatelier's Principle.
3. Explain the terms exothermic and endothermic.
4. Make sure the students relate a saturated solution to an equilibrium system.

Preparation of Solutions:

Saturated NaCl: Add 36 g of NaCl to 100 mL of water and stir for 10 minutes. Some undissolved NaCl may remain.

Saturated $KNO_3$: Add 50 g of $KNO_3$ to 100 mL of water and stir for 10 minutes. Allow any undissolved $KNO_3$ to settle to the bottom.

Distilled or deionized water is not required for these solutions.

Sample Data:

(A) A precipitate of NaCl(s) should form.
(B) a. The solution becomes cooler, this indicates an endothermic process.
    b. A precipitate of $KNO_3(s)$ should form.

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Applying "Stress" To A System In Equilibrium, page 2
Teacher's Guide

Answers to Questions:

1. The concentrated HCl has a much higher $Cl^-$ content than the saturated NaCl, thus the equilibrium will be shifted to the left. A precipitate of NaCl will result.

2. The process is endothermic. The solution has cooled as the $KNO_3$ dissolves. This means that the process absorbed energy from the surroundings which is our definition of endothermic.

3. $KNO_3(s) \rightleftharpoons K^+(aq) + NO_3^-(aq)$

   $KNO_3(s) + \underline{\underline{\text{Heat}}} \rightleftharpoons K^+(aq) + NO_3^-$

4. "Heat" in an equilibrium system can be treated as a component of the system. Thus if "heat" is removed (cooling the system) in #3, the equilibrium shifts to the left. This forms a precipitate of $KNO_3(s)$.

Waste Disposal: All of these solutions may be flushed down the drain with a large amount of running water.

# EXPERIMENT 15

# OXIDATION-REDUCTION REACTIONS OF METALS AND METAL IONS

## GENERAL TOPIC: XVI. OXIDATION-REDUCTION

**Introduction:** Metals can act as reducing agents, which means that they can react with other substances such as metallic ions, and in the process electrons are transferred. For example, copper reacts with silver ions, $Ag^+(aq)$, in aqueous solution to produce silver metal and copper ions.

$$Cu(s) \quad + \quad 2\ Ag^+(aq) \quad \text{-->} \quad 2\ Ag(s) \quad + \quad Cu^{+2}(aq)$$

The copper atoms transfer electrons to the silver ions. Copper ions go into solution and silver metal deposits out of solution. This is an oxidation reduction reaction. The copper atoms reduce the silver ions and the copper is oxidized in the process. The reverse reaction:

$$2\ Ag(s) \quad + \quad Cu^{+2}(aq) \quad \text{-->} \quad 2\ Ag^+(aq) \quad + \quad Cu(s)$$

does not occur spontaneously. This indicates that copper is a better reducing agent than silver.

**Objective:** In this experiment we shall compare the reducing power of five metals. This can be done by observing the reactions of these metals with solutions of their ions.

**Materials and Equipment:**

paper for labeling

9 test tubes (15 mm x 125 mm or larger)

test tube rack

**Chemicals:**

four pieces of copper metal

0.2 M aqueous solution of copper(II) nitrate, $Cu(NO_3)_2$

four pieces of iron metal

0.2 M aqueous solution of iron(II) sulfate, $FeSO_4$

four pieces of lead metal

0.2 M aqueous solution of lead(II) nitrate, $Pb(NO_3)_2$

four pieces of magnesium metal

0.2 M aqueous solution of magnesium nitrate, $Mg(NO_3)_2$

four pieces of zinc metal

0.2 M aqueous solution of zinc nitrate, $Zn(NO_3)_2$

distilled or deionized water (for use in preparation of solutions)

Experiment 15
Oxidation-Reduction Reactions of Metals and Metal Ions, page 2

Procedure: You have available pieces of metallic Cu, Fe, Pb, Mg, and Zn. The cations of these metals are available in the aqueous solutions of the compounds listed below.

$Cu(NO_3)_2$, $FeSO_4$, $Mg(NO_3)_2$, $Pb(NO_3)_2$ and $Zn(NO_3)_2$

Investigate what occurs when samples of each of the metallic elements are mixed with solutions of the other four metallic ions.

1. Obtain 4 pieces of each of the five metals. Place these on a sheet of paper that clearly identifies each metal.
2. Using a graduated cylinder measure approximately 45 mL of the 0.2 M $Cu(NO_3)_2$ solution and pour into a small beaker to take to your station. This is your source of $Cu^{+2}$ ions.
3. Measure 10 mL of water into one of your test tubes to use as a visual guide to fill each test tube. Fill each of 4 test tubes with 10 mL of 0.2 M $Cu(NO_3)_2$.
4. Identify the metal to be placed in each test tube by placing a labeled sheet of paper in front of the test tube rack. Use the order as shown in Table 1 on the next page. Do not put copper metal in the $Cu^{+2}(aq)$ solution.
5. Place each metal in its appropriate test tube.
6. Observe and record in Table 1 any changes which occur. Allow at least 5 minutes for observations.

   Particularly notice whether the following changes occur:

   (a) Formation of a grey or black deposit on the metal indicating the formation of a metallic form of an element.

   (b) A change in the color or appearance of the solution.

   (c) The metal appears to go into solution by "dissolving".

Experiment 15
Oxidation-Reduction Reactions of Metals and Metal Ions, page 3

7. Go back to step 2 and repeat these procedures using the next solution ($Fe^{+2}$) as shown in Table 1. During the five minute observation period you should clean your previously used test tubes as directed by your instructor. Do not pour solid pieces of metal into the sink.

8. Repeat the procedure with the other three solutions.

Data and Results:

TABLE 1

| | Cu | Fe | Mg | Pb | Zn |
|---|---|---|---|---|---|
| $Cu^{+2}$ | X | | | | |
| $Fe^{+2}$ | | X | | | |
| $Mg^{+2}$ | | | X | | |
| $Pb^{+2}$ | | | | X | |
| $Zn^{+2}$ | | | | | X |

Conclusions: 1.(a) Describe what you observed when you placed Mg metal in the $Cu(NO_3)_2$ solution.

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Oxidation-Reduction Reactions of Metals and Metal Ions, page 4

(b) Try to give an explanation in terms of a chemical reaction for what was observed .

2.(a) Describe what you observed when you placed Cu metal in the $Mg(NO_3)_2$ solution.

(b) Give an explanation for this.

3.(a) From your observations described in your answers to questions 1 and 2, what conclusion can you draw about the relative reducing power of Mg and Cu?

(b) Explain.

4.(a) Which element, of the five tested, appears to be the best reducing agent?

(b) Give a reason for your answer.

5.(a) List from top to bottom the five metals according to their strength as reducing agents, from the strongest (top) to the weakest (bottom).

Experiment 15
Oxidation-Reduction Reactions of Metals and Metal Ions, page 5

(b) Compare your listing with that of the activity series found in your text or given to you by your teacher.

Further Study: Use the information gained in this experiment to identify the metal ion present in an unknown solution. The solution will contain one of the following ions:

$Fe^{+2}$, $Mg^{+2}$, $Pb^{+2}$ or $Zn^{+2}$.

All of the materials and chemicals listed on page 1 will be available. Record your observations and conclusions.

## EXPERIMENT 15 -- TEACHER'S GUIDE

## OXIDATION-REDUCTION REACTIONS OF METALS AND METAL IONS

## GENERAL TOPIC: XVI. OXIDATION-REDUCTION

**Introduction:** This experiment illustrates oxidation-reduction reactions and the activity series of metals. If metal A reacts with the ion of another metal B then metal A is the stronger reducing agent. This is illustrated by the reaction of Cu metal and $Ag^+$ ion presented on page 1 of the experiment. This experiment also gives the student the opportunity to observe an element in two different oxidation states; for example, copper as a metal in the zero (0) oxidation state and as an ion in solution in the +2 oxidation state, etc.

**Objectives:** The student should be able to:

1. make observations and determine whether a reaction occurs between a particular metal and a metal ion in solution.
2. relate the observed change to a chemical equation as in question 1.
3. determine the order of reducing strength (activity series) by organizing the data and making comparisons.
4. design a procedure for identifying an unknown metal ion using the results of the experiment. (optional)

**Pre-laboratory Discussion:** The teacher may want to discuss oxidation states and also the definition of oxidation and reduction. He/she could use the Cu and $Ag^+$ reaction to illustrate these principles. A demonstration of this reaction would also be helpful. For example, the addition of a piece of Cu wire to a 0.5 M solution of $AgNO_3$ results in the formation of Ag crystals on the copper wire and eventually the colorless solution acquires a blue color.

**Preparation:** Required for each student or pair of students:

1. Four pieces of each metal: Fe, Cu, Mg, Pb, and Zn. Pieces should be approximately 3 cm x 1 cm strips of metal made from sheet or ribbon metal or a 3 cm piece of heavy wire. Mossy Zn can be used and a nail may be used for Fe. If all five metals are not available, the experiment can be done with four metals.

Experiment 15
Oxidation-Reduction Reactions of Metals and Metal Ions, page 2
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2. 50 mL of each solution listed below per student or pair of students:

| | |
|---|---|
| 0.2 M $Cu(NO_3)_2$ | 48.4 g of $Cu(NO_3)_2 \cdot 3H_2O$ per liter of solution |
| 0.2 M $FeSO_4$ | 55.4 g of $FeSO_4 \cdot 7H_2O$ per liter of solution |
| 0.2 M $Mg(NO_3)_2$ | 51 g of $Mg(NO_2)_2 \cdot 6H_2O$ per liter of solution |
| 0.2 M $Pb(NO_3)_2$ | 67 g of $Pb(NO_3)_2$ per liter of solution |
| 0.2 M $Zn(NO_3)_2$ | 59 g of $Zn(NO_3)_2 \cdot 6H_2O$ per liter of solution |

Other soluble salts of the metals can be used, e.g. sulfates, chlorides; however, the color change occurring may be different.

Sample Data
Conclusions:

| | Cu | Fe | Mg | Pb | Zn |
|---|---|---|---|---|---|
| $Cu^{+2}$ | X | dark deposit on Fe, solution gets cloudy | Mg darkens, solution turns green | Pb darkens, solution turns green | Red solid forms, solution turns green |
| $Fe^{+2}$ | No Reaction | X | Mg darkens solution gets cloudy | No Reaction | Zn darkens |
| $Mg^{+2}$ | No Reaction | No Reaction | X | No Reaction | No Reaction |
| $Pb^{+2}$ | No Reaction | Fe darkens | Mg reacts solution gets cloudy | X | Zn reacts |
| $Zn^{+2}$ | No Reaction | No Reaction | Mg darkens | No Reaction | X |

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Oxidation-Reduction Reactions of Metals and Metal Ions, page 3
Teacher's Guide

**Answers to Questions:** Answers to questions asked in "Conclusions" are given below:

1. A red deposit formed in the solution and the Mg dissolved. The solution also turned green. The equation is:

$$Mg(s) + Cu^{+2}(aq) \rightarrow Mg^{+2}(aq) + Cu(s)$$

2. There was no reaction. This indicates that Cu metal cannot reduce $Mg^{+2}$

3. Mg is a better reducing agent than Cu. Mg reduced $Cu^{+2}$ but Cu cannot reduce $Mg^{+2}$

4. Mg reacted with all of the solutions. Therefore it is the best reducing agent.

5. $Mg > Zn > Fe > Pb > Cu$

**Further Study:** The procedure for identifying the metal ion would be to test it with the metals and observe the changes. Compare the results with those in the table. Copper ion is not included because it would be obvious from its blue color.

**Post-laboratory Discussion:** The teacher may want to have students compare results and conclusions to see if there are any differences among students.

The teacher could also point out that the metal ions are oxidizing agents and ask students to determine the relative order of metal ions as oxidizing agents.

A related demonstration which could be done by the teacher would be to show where $H_2$ comes in the activity series by reacting the metals with an acid solution such as hydrochloric acid and observing the release of $H_2$.

**Waste Disposal:** Metal pieces should be removed from the solutions by decanting the solution from the metal or removing the metal with forceps. The metal pieces should be rinsed with water to remove drops of the solutions from which they came. Affected metal pieces should be discarded with solid waste. Unreacted metals can be reused. All solutions may be washed down the drain with a large amount of running water except for the $Pb(NO_3)_2$ solution. Dispose of it by the procedure described under "General Method for Disposing of Solutions of Toxic Inorganic Salts" found near the end of Chapter 5.

## EXPERIMENT 16

## ELECTROLYSIS OF POTASSIUM IODIDE IN WATER

### GENERAL TOPIC: XVII ELECTROCHEMISTRY

Introduction: The conversion of chemical energy to electricity, the induction of a chemical reaction, and the conductivity of many ionic solutions can be understood in terms of oxidation and reduction processes taking place in water solution. Electrochemistry is a subject which relates the chemical process to the transfer of electrons. When an atom or molecule gains electrons it is said to be reduced. The gain of electrons is called a reduction process and the reverse process, loss of electrons, is called oxidation. Whenever electrons are lost from one species, they must be gained by another species; oxidation and reduction must be accompanied by each other.

Objective: In this experiment electrical energy will be introduced into a solution containing cations and anions, and the resulting chemical changes will be observed.

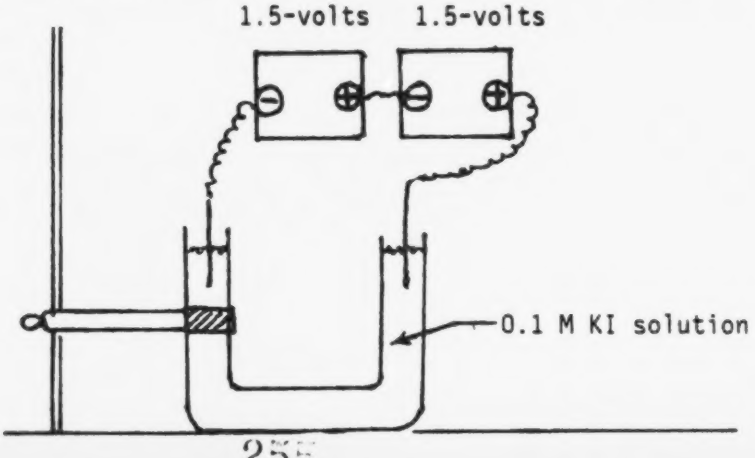
Materials and Equipment:

- two 1.5-volt lantern batteries or one 9-volt battery
- buret stand with a clamp
- litmus paper or pH paper.
- U-tube
- two 4-inch copper wires (13 to 18 gauge)

Chemicals: 0.1 M aqueous solution of potassium iodide, KI
(A sodium iodide solution could be substituted.)

Procedure:

1. Clamp a U-tube to a buret stand and connnect two 1.5-volt batteries with a copper connecting wire and with two copper "electrode" wires as shown in the drawing below.



Experiment 16
Electrolysis of Potassium Iodide, page 2

2. Pour a 0.1 M KI solution into the U-tube carefully until the two wires are immersed in the solution to a depth of 1 cm.

3. Observe carefully the process that occurs, and record all of your observations.

Results: What you observed in that arm of the U-tube which is in contact with the positive terminal of the battery is the oxidation of iodide ions in the solution. The negatively charged iodide ions (anions) move toward the positive terminal which is also called the anode. The iodide ions release electrons (undergo oxidation) at the anode and become solid iodine particles. The iodine solid, however, is not very soluble in water but will react easily with iodide ions present in water to form soluble triiodide ions having a dark brown color.

If something is being oxidized in the anode compartment, something must be undergoing reduction in the cathode compartment (since oxidation of some species must be accompanied by reduction of some other species). In the other arm of the U-tube positively charged potassium ions (cations) migrate to the negative terminal called the cathode. The potassium ions are very stable in water; therefore, no reaction (reduction) of the potassium ions will take place. The substance which undergoes reduction is water. As a result of the reduction of water, hydrogen gas is generated.

Questions:

1. Write chemical equations to show what is happening at the positive terminal in the U-tube.

2. Describe, with an equation, the generation of hydrogen gas in the cathode section of the U-tube.

3. As a consequence of the reaction, what happens to the solution? Test the solution with both red and blue litmus paper or with pH paper. Explain any color change you may see in the paper.

# EXPERIMENT 16 - TEACHER'S GUIDE

## ELECTROLYSIS OF POTASSIUM IODIDE IN WATER

## GENERAL TOPIC: XVII. ELECTROCHEMISTRY

Objectives: To show students:

1. that electric charges can induce a chemical reaction, (an oxidation-reduction reaction), and that a battery is a source of electric charges.
2. the construction of a simple electrolytic cell and an electrolysis circuit.
3. that cations in a solution of an electrolyte migrate to the cathode (negatively charged terminal), while anions move toward the anode (positively charged terminal).
4. that iodide ions (anions) migrating to the anode are oxidized to iodine molecules which react with iodide ions present in the vicinity to form triiodide ions ($I_3^-$).
5. that potassium cations migrating to the cathode are not reduced to potassium metal because these ions are very stable in solution. Water molecules have a greater tendency to be reduced than potassium ions do.
6. that the solution near the cathode will become basic as the result of hydroxide ions produced by the reduction of water and that bubbles of hydrogen gas will be generated also.
7. that the solution becomes basic (indicator used).

Pre-laboratory Discussion: Some oxidation-reduction reactions do not occur spontaneously. However, such non-spontaneous reactions can be forced to occur by the help of electric charges or current supplied externally. A cell for electrolysis (electrolytic cell) consists of a pair of electrodes made of metal wires (or carbon rods) and a solution of an electrolyte in a glass container. An electric current supplied by a battery (direct current source) is connected across the cell. One electrode which is connected to the negative terminal is called the cathode; the other electrode which is connected to the positive terminal is called the anode. When water undergoes electrolysis, oxidation of water takes place at the anode producing oxygen gas, while reduction at the cathode produces hydrogen gas.

Electrolysis of some aqueous salt solutions, however, may lead to the oxidation or reduction of the ions from the salt if these ions are more easily oxidized or reduced than water itself. It depends on the reduction (or oxidation) potential of metal ions with respect to the reduction (or oxidation) potential of water. In this experiment an aqueous solution of potassium iodide will be electrolyzed and the products formed will be identified.

The oxidation of water at the anode:

$$2\ H_2O \longrightarrow 4\ H^+(aq) + O_2(g) + 4\ e^-$$

or, if the hydronium ion is to be shown,

$$6\ H_2O \longrightarrow 4\ H_3O^+(aq) + O_2(g) + 4\ e^-$$

The reduction of water at the cathode:

$$2\ H_2O + 2\ e^- \longrightarrow H_2(g) + 2\ OH^-(aq)$$

The overall reaction between the anode and the cathode:

The anode and the cathode reactions are combined through the least common multiple of 4 and 2, that is 4.

$$2\ H_2O \longrightarrow 4\ H^+(aq) + O_2(g) + 4e^-$$

$$4\ H_2O + 4e^- \longrightarrow 2\ H_2(g) + 4\ OH^-(aq)$$

$$\overline{6\ H_2O + 4\ e^- \longrightarrow 4\ H^+(aq) + O_2(g) + 2\ H_2(g) + 4\ OH^-(aq) + 4\ e^-}$$

or

$$2\ H_2O \longrightarrow O_2(g) + 2\ H_2(g)$$

Potassium iodide in water is completely dissociated to potassium ions and iodide ions:

$$KI(s) + H_2O \longrightarrow K^+(aq) + I^-(aq)$$

The negatively charged iodide ions (anions) migrate toward the anode (+) where they are oxidized to $I_2(s)$ but the solid form of iodine is not readily soluble in water. However, when iodine molecules are surrounded by iodide ions they react to form complex triiodide ions, which dissolve in water.

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Electrolysis of Potassium Iodide in Water, page 3
Teacher's Guide

$$2\ I^-(aq) \rightarrow I_2(s) + 2\ e^- \quad \text{(oxidation of } I^-\text{)}$$

$$I_2(s) + I^-(aq) \rightarrow I_3^-(aq) \quad \text{(formation of } I_3^-\text{)}$$

Potassium ions (cations) in solution migrate toward the cathode (-), but the reduction of potassium ions in water will not take place since these ions are very stable in water. The reduction potential (tendency to be reduced) of potassium ions is much smaller than the reduction potential of water. Since water molecules have higher reduction potential than potassium ions, water molecules undergo the reduction as described above to form hydrogen gas, which is observed as gas bubbles.

$$2\ H_2O + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$$

As the result of the reduction of water at the cathode the solution in the cathode area becomes basic. This can be tested with pH paper, litmus paper or a drop of phenolphthalein solution. Red litmus will change its color to blue; with the addition of phenolphthalein, the solution will become bright pink in color.

**Preparation of Solution:** Weigh about 3.5 grams of potassium iodide (or sodium iodide) on a piece of wax paper, transfer into a 250 mL beaker, and add 200 mL of distilled or deionized water with stirring to dissolve every particle of solid. Assemble batteries and wires as shown in the figure. If the formation of triiodide ion is slow, use three 1.5-volt lantern batteries in series:

**Answers to Questions:**

1. Positive terminal (anode): the oxidation of $I^-$ ions,

$$2\ I^-(aq) \rightarrow I_2(s) + 2\ e^-$$

$$I_2(s) + I^- \rightarrow I_3^-(aq)$$

The overall process is:

$$3\ I^-(aq) \rightarrow I_3^-(aq) + 2\ e^-$$

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Electrolysis of Potassium Iodide in Water, page 4
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2. Negative terminal (cathode): the reduction of water (or hydrogen ions),

$$2\ H_2O + 2\ e^- \dashrightarrow H_2(g) + 2\ OH\text{-}(aq)$$

$$[\text{or } 2\ H^+(aq) + 2\ e^- \dashrightarrow H_2(g)]$$

3. The solution in the cathode compartment becomes basic due to the hydroxide ions generated from the reduction of water. pH paper should indicate color in the basic region; red litmus paper turns blue. A drop of phenolphthlein will produce a pink color in the solution.

Waste Disposal: After disconnecting wires from the battery and removing them from the solution, the solution can be flushed down a sink drain with plenty of running water. The wires can be rinsed, wiped dry, and used again.

# EXPERIMENT 17

## ELECTROPLATING COPPER

### GENERAL TOPIC: XVII. ELECTROCHEMISTRY

**Introduction:** Depositing or bonding a metallic layer to another substance is called electroplating. Electroplating allows such metals as silver, gold, copper, nickel, zinc or chromium to cover the surface of another metal such as iron. Electroplating is usually done either to enhance beauty or to reduce corrosion. Inexpensive jewelry is produced by plating silver or gold on a common metal (iron or copper). After a short time, a cheap "gold" ring loses its luster as the thin layer of gold wears away to reveal the copper or iron below. Tin cans or galvanized metal plates are other examples of metal coating. A thin layer of tin is plated on steel sheets to prevent the corrosion of steel.

**Objective:** To copper plate a metallic object

**Discussion:** When a solution containing copper ions is in contact with a direct current source (a DC battery) through a copper wire and another metal object, copper metal will deposit on the other object. In this experiment a paper clip will be coated with copper. You will even be able to measure the amount of copper coated if you run the plating process long enough and if you have an analytical balance capable of weighing the very small amount of copper that deposits.

**Materials and Equipment:**

two 1.5-V DC batteries connected in series

250 mL beaker

one paper clip (or any other small metallic object)

fine sand paper

two copper wires (one to be used to connect the battery to the paper clip; the other to be used as an electrode)

**Chemicals:** 200 mL of a saturated copper sulfate solution containing 5 mL of concentrated sulfuric acid solution and 5 mL of ethanol. (BE CAREFUL WITH THIS ACIDIC SOLUTION!)

**Time Required:** 20 minutes (or 2-3 hours for quantitative measurement of the copper plated)

Experiment 17
Electroplating Copper, page 2

**Procedure:**

1. Clean a paper clip (or any other metal object) and polish it with a piece of fine sand paper.

   If you plan to determine the amount of copper plated, weigh the paper clip on an analytical balance. Also weigh the copper wire to be used as an electrode.

2. Connect the paper clip to the negative terminal of one battery with a connecting wire and fasten a copper wire to the positive terminal of the other. (See Figure 1)

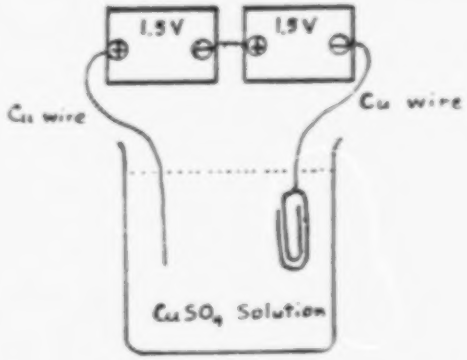
3. Immerse the paper clip and the copper wire into the copper sulfate solution contained in a 250 mL beaker. Make sure that the paper clip is covered with the solution. (BE CAREFUL!) Observe what is happening to the paper clip and the wire.

4. After about 20 minutes (or 2 to 3 hours for the quantitative experiment) disconnect wires from the batteries and rinse both wires gently with water. Observe both the paper clip and the copper wire to see any change that took place.

   If you ran the electrolysis for more than 2 hours, weigh the dried paper clip and find how much its mass has changed. Also weigh the copper wire and determine how much its mass has changed.

**Results and Conclusions:** The paper clip, connected by wire to the negative terminal of the battery, is called the cathode of this electrolysis cell. Copper(II) ions will migrate toward the paper clip and will be reduced to copper metal. An oxidation process will take place at the copper wire fastened to the positive terminal. Copper metal from this wire will be converted into copper(II) ions and electrons.

Fig. 1



**Problems:**

1. Write the chemical equation for the reaction at the anode.

Experiment 17
Electroplating Copper, page 3

2. Write the chemical reaction equation for the process taking place at the cathode.

3. Explain why water isn't converted to $H_2$ and $OH^-(aq)$ by reduction.

4. Describe any visible change that occurs in the wire, paper clip, or solution.

# EXPERIMENT 17 - TEACHER'S GUIDE

# ELECTROPLATING COPPER

## GENERAL TOPIC: XVII. ELECTROCHEMISTRY

**Objectives:** Students will learn:

1. to electroplate a metallic object.
2. the solution chemistry taking place during an electroplating process.
3. the application of electrochemistry (an oxidation-reduction process) to achieve a useful objective.

**Pre-laboratory Discussion:** In water solution ions of metals low in the activity series are more easily reduced to the metals than are hydrogen ions to hydrogen gas. This fact makes possible an electrolytic process called electroplating.

An electroplating cell contains a solution of a compound that furnishes ions of the metal to be plated at the cathode and uses a piece of the plating metal as the anode. A copper plating cell, for example, contains a solution of a soluble copper compound and copper metal as the anode. The cathode is the object to be plated. As shown in the figure, the copper wire is connected to the positive terminal of a battery and the metal object to be plated is connected to the negative terminal.

Electrons are supplied to the cathode of the electroplating cell from the battery, and copper ions in solution gain electrons at the cathode. The electron-gaining process by copper ions is an example of reduction.

$$Cu^{+2}(aq) + 2e^- \dashrightarrow Cu(s)$$

As the copper ions gain electrons, copper metal will be deposited on the metal object.

Meanwhile at the anode, oxidation of copper metal will take place.

$$Cu(s) \dashrightarrow Cu^{+2}(aq) + 2e^-$$

Experiment 17
Electroplating Copper, page 2
Teacher's Guide

Electrons are released (or lost) from the copper metal and flow through the external circuit. Copper ions generated at the anode will migrate to the cathode through the solution. Thus, in effect, copper is transferred from the anode to the cathode of the cell.

**Preparation of Solution:**

1. Saturated copper(II) sulfate solution: Weigh out 62 grams of $CuSO_4 \cdot 5H_2O$ and dissolve it in 200 mL of either distilled or deionized water at room temperature. If it takes too long for the solid to dissolve, heat the solution gently on a hot plate; or, if you leave the mixture overnight, the solid will dissolve.

2. Add about 5 mL of concentrated sulfuric acid (or 10-20 mL of dilute sulfuric acid solution) to the copper(II) sulfate solution slowly with stirring. CAREFUL! SINCE SULFURIC ACID IS A STRONGLY CORROSIVE ACID, GREAT CARE SHOULD BE TAKEN. DO NOT LET A STUDENT PREPARE THIS SOLUTION. DO NOT ADD COPPER(II) SULFATE SOLUTION TO ACID.

3. Add about 5 mL of ethyl alcohol to the solution.

4. The reason for adding some sulfuric acid and ethyl alcohol to the copper sulfate solution is to increase the conductivity of the solution and to make the plating smooth and clean.

5. Connect two batteries as shown in the figure, and attach copper wires and a metal object to be plated as indicated in the figure.

**Answers to Problems:**

1. The anode reaction (positive terminal): The oxidation of copper atoms to copper ions and the release of electrons.

   $$Cu(s) \longrightarrow Cu^{+2}(aq) + 2\ e^-$$

2. The cathode reaction (negative terminal): the reduction of copper ions to copper metal resulting in the deposition of copper on a metal object.

   $$Cu^{+2}(aq) + 2\ e^- \longrightarrow Cu(s)$$

3. The reduction of water at the cathode can be expressed by the equation:

$$2\ H_2O\ +\ 2\ e^-\ \dashrightarrow\ H_2(g)\ +\ 2\ OH^-(aq)$$

The tendency of water to be reduced to $H_2(g)$ and $OH^-(aq)$ is much less than that of copper ion in water to be reduced to copper metal. In other words, the reduction potential of $Cu^{+2}(aq)$ is greater than that of $H_2O$. As long as copper ions remain in the solution, only a very small amount of water will be reduced to hydrogen.

4. The paper clip becomes coated with copper.

**Post-laboratory Discussion:** Explain the anode-cathode reaction and the oxidation-reduction process. Make the association between the oxidation-reduction process and the anode-cathode reaction.

If you have an analytical balance with the sensitivity of 0.0001 g (0.1 mg), you can determine the mass of copper deposited on the metal object.

mass of copper deposited = mass of metal object after electroplating - mass of metal object before electroplating

**Waste Disposal:** Dispose of the copper sulfate solution in a sink with a large amount of running water or you can pour the solution into a bottle for later use.

# CHAPTER 9

# REFERENCE MATERIAL

# CHAPTER 9 -- REFERENCE MATERIAL

## UNITS USED THROUGHOUT THIS BOOK

| Units | Abreviation | Alternate Units | Abreviation |
|---|---|---|---|
| **Mass** | | | |
| gram | g | | |
| kilogram | kg | | |
| milligram | mg | | |
| **Length** | | | |
| meter (metre) | m | | |
| centimeter | cm | | |
| millimeter | mm | | |
| **Volume** | | | |
| liter (litre) | L | cubic decimeter | $dm^3$ |
| milliliter | mL | cubic centimeter | $cm^3$ |
| **Pressure** | | | |
| millimeters of mercury | mm Hg | torr | torr |
| atmosphere | atm. | pascal (conversion number) | Pa |
| **Temperature** | | | |
| Celsius | °C | | |
| Kelvin | K | | |
| **Energy** | | | |
| Joules | J | calorie (1 cal = 4.184J) | cal |
| **Concentration** | | | |
| molarity (moles per liter) | M | moles per cubic decimeter | $mol/dm^3$ |

## MOLECULAR MASSES OF COMMONLY USED COMPOUNDS

| | | | |
|---|---|---|---|
| $AgBr$ | 187.78 | $CuO$ | 79.54 |
| $AgCl$ | 143.32 | $Cu_2O$ | 143.08 |
| $AgI$ | 234.77 | $CuSO_4 \cdot 5H_2O$ | 249.68 |
| $AgNO_3$ | 169.88 | $CuS$ | 95.60 |
| $Al_2O_3$ | 101.96 | $FeO$ | 71.85 |
| $Al(OH)_3$ | 78.00 | $Fe_2O_3$ | 159.69 |
| $Al_2(SO_4)_3$ | 342.15 | $Fe_3O_4$ | 231.54 |
| $BaBr_2$ | 197.16 | $Fe(OH)_2$ | 89.87 |
| $BaCO_3$ | 197.35 | $Fe(OH)_3$ | 106.87 |
| $BaCl_2$ | 208.24 | $FeS$ | 87.91 |
| $BaCrO_4$ | 253.34 | $FeSO_4 \cdot 7H_2O$ | 278.02 |
| $BaF_2$ | 175.34 | $Fe_2(SO_4)_3$ | 399.89 |
| $Ba(OH)_2$ | 171.36 | $H_3BO_3$ | 61.83 |
| $Ba_3(PO_4)_2$ | 601.97 | $HBr$ | 80.92 |
| $BaSO_4$ | 233.40 | $HC_2H_3O_2$ (acetic) | 60.05 |
| $CaBr_2$ | 199.90 | $HCl$ | 36.46 |
| $CaCO_3$ | 100.09 | $HClO_4$ | 100.46 |
| $CaF_2$ | 78.08 | $HF$ | 20.01 |
| $CaO$ | 56.08 | $HNO_2$ | 47.01 |
| $Ca(OH)_2$ | 74.10 | $HNO_3$ | 63.02 |
| $Ca(NO_3)_2$ | 164.10 | $H_2O_2$ | 34.02 |
| $Ca_3(PO_4)_2$ | 310.19 | $H_3PO_4$ | 97.99 |
| $CaSO_4$ | 136.14 | $H_2S$ | 34.08 |
| $CO$ | 28.01 | $H_2SO_3$ | 82.08 |
| $CO_2$ | 44.01 | $H_2SO_4$ | 98.08 |
| $Cu(NO_3)_2$ | 187.54 | $HgO$ | 216.59 |

(Continued on next page)

## Molecular Masses of Commonly Used Compounds, page 2

| | | | |
|---|---|---|---|
| HgS | 232.65 | $Mg(ClO_4)_2$ | 223.22 |
| $Hg_2Br_2$ | 561.00 | $MgI_2$ | 278.12 |
| $Hg_2Cl_2$ | 472.08 | MgO | 40.31 |
| $Hg_2I_2$ | 654.99 | $Mg(OH)_2$ | 58.33 |
| KBr | 119.01 | $MgSO_4$ | 120.37 |
| $KBrO_3$ | 167.01 | $MnO_2$ | 86.94 |
| KCN | 65.12 | MnS | 87.00 |
| KCl | 74.56 | NaBr | 102.90 |
| $KClO_3$ | 122.55 | $NaC_2H_3O_2$ (acetate) | 82.03 |
| $KClO_4$ | 138.55 | NaCN | 49.01 |
| $K_2CrO_4$ | 194.20 | NaCNS | 81.07 |
| $K_2Cr_2O_7$ | 294.20 | $Na_2CO_3$ | 105.99 |
| KI | 166.00 | NaCl | 58.44 |
| $KIO_3$ | 214.00 | NaClO | 74.44 |
| $KIO_4$ | 230.00 | $NaClO_2$ | 90.44 |
| $KMnO_4$ | 158.04 | $NaHCO_3$ | 84.01 |
| $KNO_3$ | 101.11 | $NaH_2PO_4$ | 119.99 |
| KOH | 56.11 | $Na_2HPO_4$ | 141.98 |
| $K_3PO_4$ | 212.01 | NaI | 149.89 |
| $K_2SO_4$ | 174.26 | $NaNO_2$ | 69.00 |
| $K_2S_2O_8$ | 270.32 | $NaNO_3$ | 85.00 |
| LiBr | 86.86 | $Na_2O$ | 61.98 |
| LiCl | 42.39 | $Na_2O_2$ | 77.98 |
| $Li_2SO_4$ | 109.94 | NaOH | 40.00 |
| $MgCO_3$ | 84.32 | $Na_3PO_4$ | 163.95 |
| $MgCl_2$ | 95.22 | $Na_2S$ | 78.04 |

(Continued on next page)

Molecular Masses of Commonly Used Compounds, page 3

| | | | |
|---|---|---|---|
| $Na_2SO_3$ | 126.04 | $SnO_2$ | 150.69 |
| $Na_2SO_4$ | 142.04 | SnS | 150.75 |
| $NH_3$ | 17.03 | $SO_2$ | 64.06 |
| $(NH_4)_2CO_3$ | 96.09 | $SO_3$ | 80.06 |
| $NH_4Cl$ | 53.49 | $SrCO_3$ | 147.63 |
| $(NH_4)_2SO_4$ | 132.13 | $Sr_3(PO_4)_2$ | 452.81 |
| $NiCl_2$ | 129.62 | $SrSO_4$ | 183.68 |
| $P_2O_5$ | 141.95 | $TiO_2$ | 79.90 |
| $PbCrO_4$ | 323.19 | $U_3O_8$ | 842.09 |
| $PbF_2$ | 245.19 | ZnS | 97.43 |
| $PbI_2$ | 461.00 | $ZnSO_4$ | 161.43 |
| PbO | 223.19 | Water for hydrates: | |
| $PbO_2$ | 239.19 | 1 $H_2O$ | 18.02 |
| $Pb(OH)_2$ | 241.21 | 2 $H_2O$ | 36.03 |
| PbS | 239.25 | 3 $H_2O$ | 54.05 |
| $PbSO_4$ | 303.25 | 4 $H_2O$ | 72.06 |
| $SiF_4$ | 104.08 | 5 $H_2O$ | 90.08 |
| $SiO_2$ | 60.09 | 6 $H_2O$ | 108.09 |
| $SnCl_2$ | 189.60 | 7 $H_2O$ | 126.11 |

## NAMES AND CHARGES OF IMPORTANT CATIONS

| Name | Formula and Charge |
|---|---|
| Aluminum | $Al^{+3}$ |
| Ammonium | $NH_4^+$ |
| Barium | $Ba^{+2}$ |
| Cadmium | $Cd^{+2}$ |
| Calcium | $Ca^{+2}$ |
| Chromium(III), chromic | $Cr^{+3}$ |
| Cobalt(II), cobaltous | $Co^{+2}$ |
| Copper(II), cupric | $Cu^{+2}$ |
| Hydrogen, hydronium | $H^+$, $H_3O^+$ |
| Iron(II), ferrous | $Fe^{+2}$ |
| Iron(III), ferric | $Fe^{+3}$ |
| Lead | $Pb^{+2}$ |
| Lithium | $Li^+$ |
| Magnesium | $Mg^{+2}$ |
| Manganese(II), manganous | $Mn^{+2}$ |
| Mercury(I), mercurous | $Hg_2^{+2}$ |
| Mercury(II), mercuric | $Hg^{+2}$ |
| Nickel | $Ni^{+2}$ |
| Potassium | $K^+$ |
| Silver | $Ag^+$ |
| Sodium | $Na^+$ |
| Strontium | $Sr^{+2}$ |
| Tin(II), stannous | $Sn^{+2}$ |
| Zinc | $Zn^{+2}$ |

## NAMES AND CHARGES OF IMPORTANT ANIONS

| Name | Formula and Charge |
|---|---|
| Acetate, $CH_3COO^-$ | $C_2H_3O_2^-$ |
| Bromate | $BrO_3^-$ |
| Bromide | $Br^-$ |
| Bromite | $BrO_2^-$ |
| Carbonate | $CO_3^{-2}$ |
| Hydrogen carbonate, bicarbonate | $HCO_3^-$ |
| Chlorate | $ClO_3^-$ |
| Chloride | $Cl^-$ |
| Chlorite | $ClO_2^-$ |
| Chromate | $CrO_4^{-2}$ |
| Dichromate | $Cr_2O_7^{-2}$ |
| Cyanate | $OCN^-$ |
| Cyanide | $CN^-$ |
| Fluoride | $F^-$ |
| Hydroxide | $OH^-$ |
| Hypobromite | $BrO^-$ |
| Hypochlorite | $ClO^-$ |
| Hypoiodite | $IO^-$ |
| Iodate | $IO_3^-$ |
| Iodide | $I^-$ |

| Name | Formula and Charge |
|---|---|
| Nitrate | $NO_3^-$ |
| Nitrite | $NO_2^-$ |
| Oxide | $O^{-2}$ |
| Perchlorate | $ClO_4^-$ |
| Permanganate | $MnO_4^-$ |
| Peroxide | $O_2^{-2}$ |
| Phosphate | $PO_4^{-3}$ |
| Monohydrogen phosphate | $PO_4^{-2}$ |
| Dihydrogen phosphate | $H_2PO_4^-$ |
| Selenide | $Se^{-2}$ |
| Silicate | $SiO_3^{-2}$ |
| Sulfate | $SO_4^{-2}$ |
| Hydrogen sulfate, bisulfate | $HSO_4^-$ |
| Sulfide | $S^{-2}$ |
| Hydrogen sulfide, bisulfide | $HS^-$ |
| Sulfide | $S^{-2}$ |
| Hydrogen sulfite, bisulfite | $HSO_3^-$ |

## ATOMIC AND IONIC RADII OF SELECTED ELEMENTS

| Atom | Atomic Radius in Å | Ion | Ionic Radius in Å |
|---|---|---|---|
| H | 0.37 | | |
| He | 0.93 | | |
| Li | 1.52 | $Li^{+}$ | 0.60 |
| Be | 1.11 | $Be^{+2}$ | 0.31 |
| B | 0.88 | | |
| C | 0.77 | | |
| N | 0.70 | | |
| O | 0.66 | $O^{-2}$ | 1.40 |
| F | 0.64 | $F^{-}$ | 1.36 |
| Ne | 1.12 | | |
| Na | 1.84 | $Na^{+}$ | 0.95 |
| Mg | 1.60 | $Mg^{+2}$ | 0.65 |
| Al | 1.43 | $Al^{+3}$ | 0.50 |
| Si | 1.17 | | |
| P | 1.10 | | |
| S | 1.04 | $S^{-2}$ | 1.84 |
| Cl | 0.99 | $Cl^{-1}$ | 1.81 |
| Ar | 1.54 | | |
| K | 2.31 | $K^{+}$ | 1.33 |
| Ca | 1.97 | $Ca^{+2}$ | 0.99 |
| Sc | 1.60 | $Sc^{+3}$ | 0.81 |
| Ti | 1.46 | | |
| V | 1.31 | | |
| Cr | 1.25 | | |
| Mn | 1.29 | | |
| Fe | 1.26 | | |
| Co | 1.25 | | |
| Ni | 1.24 | | |
| Cu | 1.28 | | |
| Zn | 1.33 | | |
| Ga | 1.22 | | |
| Ge | 1.22 | | |
| As | 1.21 | | |
| Se | 1.17 | $Se^{-2}$ | 1.98 |
| Br | 1.14 | $Br^{-}$ | 1.95 |
| Kr | 1.69 | | |
| Rb | 2.44 | $Rb^{+}$ | 1.48 |
| Sr | 2.15 | $Sr^{+2}$ | 1.13 |
| In | 1.62 | | |
| Sn | 1.40 | | |
| Sb | 1.41 | | |
| Te | 1.37 | $Te^{-2}$ | 2.21 |
| I | 1.33 | $I^{-}$ | 2.16 |
| Xe | 1.90 | | |
| Cs | 2.62 | $Cs^{+}$ | 1.69 |
| Ba | 2.17 | $Ba^{+2}$ | 1.35 |
| Tl | 1.71 | | |
| Pb | 1.75 | | |
| Bi | 1.46 | | |
| Po | 1.65 | | |

## TRIVIAL NAMES OF CHEMICALS

| Trivial Name | Chemical Name | Formula |
|---|---|---|
| alcohol | ethyl alcohol | $CH_3CH_2OH$ |
| alum | potassium aluminum sulfate dodecahydrate | $KAl(SO_4)_2 \cdot 12H_2O$ |
| aqua regia | mixture of concentrated HCl and concentrated $HNO_3$ | $HCl + HNO_3$ |
| aspirin | acetylsalicylic acid | $C_9H_8O_4$ |
| bleaching powder | calcium chloride hypochlorite | $CaOCl_2$ |
| borax | sodium tetraborate decahydrate | $Na_2B_4O_7 \cdot 10H_2O$ |
| brimstone | sulfur | $S$ |
| calomel | mercurous chloride | $Hg_2Cl_2$ |
| carbide | calcium carbide | $CaC_2$ |
| carbolic acid | phenol | $C_6H_5OH$ |
| caustic potash | potassium hydroxide | $KOH$ |
| caustic soda | sodium hydroxide | $NaOH$ |
| chalk | calcium carbonate | $CaCO_3$ |
| dextrose | glucose | $C_6H_{12}O_6$ |
| Epsom salt | magnesium sulfate heptahydrate | $MgSO_4 \cdot 7H_2O$ |
| ether | diethyl ether | $CH_3CH_2OCH_2CH_3$ |
| fool's gold | iron disulfide | $FeS_2$ |
| freon | dichlorodifluoromethane | $CCl_2F_2$ |
| Glauber's salt | sodium sulfate decahydrate | $Na_2SO_4 \cdot 10H_2O$ |
| glycerine | glycerol | $CH_2OH\text{-}CHOH\text{-}CH_2OH$ |
| grain alcohol | ethyl alcohol | $CH_3CH_2OH$ |
| gypsum | calcium sulfate dihydrate | $CaSO_4 \cdot 2H_2O$ |
| laughing gas | nitrous oxide | $N_2O$ |
| lye | sodium hydroxide | $NaOH$ |
| marsh gas | methane | $CH_4$ |
| milk of magnesia | magnesium hydroxide suspended in water | $Mg(OH)_2$ |
| muriatic acid | impure hydrochloric acid | $HCl$ |
| peroxide | hydrogen peroxide solution | $H_2O_2$ |
| plaster of Paris | calcium sulfate half hydrate | $CaSO_4 \cdot \frac{1}{2}H_2O$ |

(Continued on next page)

Trival Names of Chemicals, page 2

| Trivial Name | Chemical Name | Formula |
|---|---|---|
| potash | potassium carbonate | $K_2CO_3$ |
| quicksilver | mercury | $Hg$ |
| Rochelle salt | potassium sodium tartrate | $KNaC_4H_4O_6$ |
| saltpeter | potassium nitrate | $KNO_3$ |
| soda ash | sodium carbonate | $Na_2CO_3$ |
| sugar | sucrose | $C_{12}H_{22}O_{11}$ |
| wood alcohol | methyl alcohol | $CH_3OH$ |

## ACTIVITY (OR REPLACEMENT) SERIES OF SOME COMMON METALS

| | | |
|---|---|---|
| Very active with $H_2O$ or acids. | K | Oxides reduced by electrolysis, but not by $H_2$ or CO. |
| | Ba | |
| | Ca | |
| | Na | |
| Active with acids, or with steam when metal is hot. | Mg | |
| | Al | |
| | Mn | Oxides reduced by C or Al (hot), but not by $H_2$ or CO. |
| | Zn | |
| | Cr | |
| | Fe | Oxides reduced by heating with $H_2$ or CO. |
| Less reactive with acids. | Cd | |
| | Ni | |
| | Sn | |
| | Pb | |
| | H | |
| Active only with stronger oxidizing acids, as $HNO_3$ and $H_2SO_4$. No $H_2$ formed. | Cu | |
| | Sb | |
| | Bi | |
| | Hg | Oxides reduced to metal (decomposed) by heat alone. |
| | Ag | |
| Active only with aqua regia | Au | |
| | Pt | |

## pH OF COMMON SUBSTANCES
(Alphabetical Listing)

| Substance | pH |
|---|---|
| baking soda | 8.4 - 9.0 |
| battery acid | 0.5 |
| beer | 4.4 - 4.5 |
| black Coffee | 5.0 |
| blood | 7.3 - 7.5 |
| eggs | 7.6 - 8.0 |
| hair remover | 12.2 - 12.8 |
| household ammonia | 10.5 - 11.9 |
| household bleach | 9.0 |
| lemon juice | 2.3 |
| milk | 6.6 |
| milk of magnesia | 9.9 - 10.5 |
| nonphosphate detergents | 11.0 - 11.5 |
| oven cleaner | 13.5 |
| phosphate detergents | 9.0 |
| pure water | 7.0 |
| saliva | 6.5 - 7.5 |
| sea water | 7.8.- 8.3 |
| shampoo | 8.0 |
| soap solutions | 10.0 |
| soft drinks | 3.9 |
| stomach (acid) | 1.0 |
| stomach (normal) | 1.0 - 3.0 |
| tomatoes | 4.0 - 4.4 |
| tums | 9.0 |
| urine | 4.5 - 8.0 |
| vinegar | 2.4 - 3.4 |
| washing soda | 12.0 |
| wine | 2.8 - 3.8 |

## pH OF COMMON SUBSTANCES
(Listed According to Increasing pH)

| Substance | pH |
|---|---|
| battery acid | 0.5 |
| stomach (acid) | 1.0 |
| stomach (normal) | 1.0 - 3.0 |
| lemon juice | 2.3 |
| vinegar | 2.4 - 3.4 |
| wine | 2.8 - 3.8 |
| soft drinks | 3.9 |
| tomatoes | 4.0 - 4.4 |
| beer | 4.4 - 4.5 |
| black coffee | 5.0 |
| urine | 4.5 - 8.0 |
| milk | 6.6 |
| saliva | 6.5 - 7.5 |
| pure water | 7.0 |
| blood | 7.3 - 7.5 |
| eggs | 7.6 - 8.0 |
| sea water | 7.8.- 8.3 |
| shampoo | 8.0 |
| baking soda | 8.4 - 9.0 |
| phosphate detergents | 9.0 |
| household bleach | 9.0 |
| tums | 9.0 |
| soap solutions | 10.0 |
| milk of magnesia | 9.9 - 10.5 |
| household ammonia | 10.5 - 11.9 |
| nonphosphate detergents | 11.0 - 11.5 |
| washing soda | 12.0 |
| hair remover | 12.2 - 12.8 |
| oven cleaner | 13.5 |

# THE COLOR CHANGES AND pH RANGE OF SOME INDICATORS

| Name of Indicator | pH Range | Color Change | Solvent |
|---|---|---|---|
| Methyl violet | 0.2-3.0 | yellow, blue, violet | water |
| Thymol blue | 1.2-2.8 | red to yellow | water (+NaOH) |
| Orange IV (tropeolin 00) | 1.3-3.0 | red to yellow | water |
| Benzopurpurin 4B | 1.2-4.0 | violet to red | 20% alcohol |
| Methyl orange | 3.1-4.4 | red to orange yellow | water |
| Bromphenol blue | 3.0-4.6 | yellow to blue violet | water (+NaOH) |
| Congo red | 3.0-5.0 | blue to red | 70% alcohol |
| Bromcresol green | 3.8-5.4 | yellow to blue | water (+NaOH) |
| Methyl red | 4.4-6.2 | red to yellow | water (+NaOH) |
| Chlorphenol red | 4.8-6.8 | yellow to red | water (+NaOH) |
| Bromcresol purple | 5.2-6.8 | yellow to purple | water (+NaOH) |
| Litmus | 4.5-8.3 | red to blue | water |
| Bromthymol blue | 6.0-7.6 | yellow to blue | water (+NaOH) |
| Phenol red | 6.8-8.2 | yellow to red | water (+NaOH) |
| Thymol blue | 8.0-9.6 | yellow to blue | water (+NaOH) |
| Phenolphthalein | 8.3-10.0 | colorless to red | 70% alcohol |
| Thymolphthalein | 9.3-10.5 | yellow to blue | 70% alcohol |
| Alizarin yellow R | 10.0-12.0 | yellow to red | 95% alcohol |
| Indigo carmine | 11.4-13.0 | blue to yellow | 50% alcohol |
| Trinitrobenzene | 12.0-14.0 | colorless to orange | 70% alcohol |

## VAPOR PRESSURE OF WATER AT VARIOUS TEMPERATURES

In mm of mercury

| Temperature °C | Pressure |
|---|---|
| 0 | 4.597 |
| 5 | 6.543 |
| 10 | 9.209 |
| 15 | 12.788 |
| 16 | 13.634 |
| 17 | 14.530 |
| 18 | 15.477 |
| 19 | 16.477 |
| 20 | 17.535 |
| 21 | 18.650 |
| 22 | 19.827 |
| 23 | 21.068 |
| 24 | 22.377 |
| 25 | 23.756 |
| 26 | 25.209 |
| 27 | 26.739 |
| 28 | 28.349 |
| 29 | 30.043 |
| 30 | 31.824 |
| 31 | 33.695 |
| 32 | 35.663 |
| 33 | 37.729 |
| 34 | 39.898 |
| 35 | 42.175 |
| 36 | 44.563 |
| 37 | 47.067 |
| 38 | 49.692 |
| 39 | 52.442 |
| 40 | 55.324 |
| 41 | 58.34 |
| 42 | 61.50 |
| 43 | 64.80 |
| 44 | 68.26 |
| 45 | 71.88 |
| 50 | 92.51 |
| 55 | 118.04 |
| 60 | 149.38 |
| 65 | 187.54 |
| 70 | 233.7 |
| 75 | 289.1 |
| 80 | 355.1 |
| 85 | 433.6 |
| 90 | 525.76 |
| 95 | 633.90 |
| 100 | 760 |

## SOLUBILITY OF GASES

1. Compounds which are gases at ordinary temperatures:

   $NH_3$, $CH_4$, HF, HCl, HBr, HI, $H_2S$, $CO_2$, $SO_2$, CO

2. Compounds which decompose at ordinary temperatures to yield gaseous products:

   $H_2CO_3 \dashrightarrow H_2O + CO_2$, $H_2SO_3 \dashrightarrow H_2O + SO_2$, $NH_4OH \dashrightarrow H_2O + NH_3$

3. Solubility of gases in water at 20°C:

   (a) Very soluble in water (400 to 700 volumes of gas per volume of water):

   HCl, HF, HBr, HI, $NH_3$

   (b) Fairly soluble in water (40 volumes of gas per volume of water):

   $SO_2$

   (c) Slightly soluble in water (from 1 to 2.5 volumes of gas per volume of water):

   $CO_2$, $Cl_2$, $H_2S$

   (d) Only very slightly soluble in water (from 1.5 to 3.0 volumes of gas per 100 volumes of water):

   $N_2$, $H_2$, CO, $O_2$, $CH_4$

## CONVERSION FACTORS

### Mass

1 gram = 0.03527 ounce (avoir.)
1 kilogram = 2.205 pounds
1 ounce (avoir.) = 28.3 grams
1 pound = 453.6 grams

### Length

1 Angstrom Unit (Å) = $10^{-8}$ centimeter
1 inch = 2.540 centimeters
1 centimeter = $1 \times 10^{8}$ Angstrom unit
1 centimeter = 0.3937 inch
1 kilometer = 0.621 miles
1 meter = 39.37 inches

### Volume

1 cubic centimeter = 0.06102 cubic inches
1 cubic inch = 16.387 cubic centimeters
1 liter = 1.0567 quarts
1 milliliter = 0.0338 ounce (U.S. liquid)
1 mole of a gas at STP = 22.413 liters
1 ounce (U.S. liquid) = 29.6 milliliters
1 quart = 946.3 milliliters
1 cubic foot = 28.3 liters
1 cubic foot = 7.48 gallons

### Electrical Energy

amperes = volts/ohms
1 ampere = 1 coulomb/second

1 faraday = 96,500 coulombs
1 faraday = $6.023 \times 10^{23}$ electrons

watts = volts x amperes

(Continued on next page)

Conversion Factors, page 2

## Energy

1 British thermal unit = 252 calories
1 calorie = energy required to raise one gram of water one degree Celsius
1 calorie = 4.184 joules
1 joule = 0.2390 calories

## Miscellaneous

Avogadro's number = $6.023 \times 10^{23}$
speed of light = $2.998 \times 10^{10}$ cm/sec

## Temperature

absolute zero ($0^{o}$K) = $-273^{o}$C
degrees Celsius = 5/9 (degrees Fahrenheit - 32)
degrees Fahrenheit = 9/5 degrees Celsius + 32
degrees Kelvin = degrees Celsius + 273

# LOCAL SOURCES OF CERTAIN COMMON CHEMICALS

| CHEMICAL NAME | COMMON NAME | SOURCE |
|---|---|---|
| acid, acetic, ca 5% | distilled vinegar | grocery store |
| acid, acetylsalicylic | aspirin | drug store |
| acid, ascorbic | vitamin C | drug store |
| acid, boric | boric acid | drug store |
| acid, citric | lemon juice | grocery store |
| acid, hydrochloric | muriatic acid | drug store |
| acid, sulfuric | battery acid | auto parts store |
| aluminum | aluminum foil | grocery store |
| ammonium chloride | sal ammoniac | drug store |
| ammonium hydroxide | ammonia | grocery store |
| calcium carbonate | chalk | drug store |
| calcium hypochlorite | calcium hypochlorite | swimming pool supply |
| calcium oxide | lime | co-op, grocery store |
| calcium sulfate | plaster of Paris | building supply |
| carbon | charcoal | grocery store |
| carbon dioxide | carbon dioxide | welding supply |
| distilled water | distilled water | grocery or drug store |
| ethyl alcohol (denatured) | ethyl alcohol | drug store |
| fructose | fructose | drug store |
| glucose | glucose | drug store |
| glycerol | glycerin | drug store |
| hydrogen peroxide (3%) | peroxide | drug store |
| iodine, in alcohol soln. | tincture of iodine | drug store |
| iron | iron nails | hardware store |
| isopropyl alcohol, 70% | rubbing alcohol | grocery or drug store |
| magnesium hydroxide suspended in $H_2O$ | milk of magnesia | grocery or drug store |
| magnesium sulfate | Epsom salts | drug store |
| mineral oil | mineral oil | drug store |
| naphthalene | moth balls (1 of 2 types) | grocery store |
| oxygen | oxygen | welding supply |
| potassium nitrate | saltpeter | drug store |
| propane | propane torch | hardware store |
| sodium chloride | salt | grocery store |
| sodium hydrogen carbonate | baking soda | grocery store |
| sodium hydroxide | lye | grocery store |
| sodium hypochlorite | bleach (Clorox) | grocery store |
| sodium phosphate $Na_3PO_4$ | TSP | co-op, hardware store |
| sodium tetraborate hydrated $Na_2B_4O_7 \cdot 10H_2O$ | Borax | grocery store |
| sucrose | sugar | grocery store |
| sulfur | flowers of sulfur | drug store |

The following substances can be found in drug stores, groceries, etc. and can be used as indicated:

Vaseline as a lubricant, Sterno as a source of heat, Alka Seltzer as a $CO_2$ source.

# LOCAL SOURCES OF CHEMISTRY SUPPLIES

| EQUIPMENT | SUGGESTED SOURCE* |
|---|---|
| bags, plastic | 3, 4, 5 |
| balances or scales | 10 |
| batteries | 3, 4, 5, 6, 7 |
| beakers | 8, 9 |
| bottles | 5, 9 |
| bottles, spray | 4, 5 |
| brass fittings | 5, 7 |
| brushes | 3, 5, 7 |
| burners, propane | 5 |
| cylinders, graduated | 5, 9 |
| files, triangular | 5, 7 |
| forceps | 5, 9 |
| funnels, metal | 5, 7 |
| funnels, plastic | 3, 5, 7 |
| hoses, plastic | 3, 5, 7 |
| hose fittings | 5, 7 |
| kits for chlorination | 5 |
| kits for testing pH | 5 |
| lighters, cigarette, disposable | 3, 4, 6 |
| masks, respiration | 5 |
| plumbing connections | 5, 7 |
| pulleys, all sizes | 5 |
| spatulas | 5 |
| stoppers, rubber | 5 |
| syringes | 8, 9 |
| syringes, large | 11 |
| test tubes | 8, 9 |
| thermometers, Celsius | 5 |
| transformers (6-12 v) | 2 |
| triangular files | 5, 7 |
| tubing, rubber | 4, 5 |
| wire, chrome | 3, 5, 7 |
| wire, electric | 3, 5, 7 |

1. air conditioning supply house
2. auto parts supply house
3. discount store
4. drug store
5. farm co-op
6. grocery store
7. hardware store
8. hospital
9. medical supply store (Jackson, MS and other larger towns)
10. State Education Agency For Surplus Property, Whitfield Road, Jackson, MS
11. veterinarian or animal hospital

*The listed equipment also might be obtained at sources other than those suggested.